Chemical Thermodynamics JEE Main PYQ - 1

Total Time: 1 Hour : 15 Minute **Total Marks:** 120

Instructions

Instructions

1. Test will auto submit when the Time is up.
2. The Test comprises of multiple choice questions (MCQ) with one or more correct answers.
3. The clock in the top right corner will display the remaining time available for you to complete the examination.

Navigating & Answering a Question

1. The answer will be saved automatically upon clicking on an option amongst the given choices of answer.
2. To deselect your chosen answer, click on the clear response button.
3. The marking scheme will be displayed for each question on the top right corner of the test window.

Thermodynamics

1. The plot of $\log K$ versus $\frac{1}{T}$ is a straight line. The intercept and slope of this line are respectively given by (Where K is the equilibrium constant.) (+4, -1)

   a. $\frac{\Delta S^\circ}{2.303R}, -\frac{\Delta H^\circ}{2.303R}$

   b. $\frac{\Delta S^\circ}{R}, -\frac{\Delta H^\circ}{R}$

   c. $-\frac{\Delta S^\circ}{2.303R}, \frac{\Delta H^\circ}{2.303R}$

   d. $-\frac{\Delta H^\circ}{2.303R}, \frac{\Delta S^\circ}{2.303R}$

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2. For an ideal gas undergo isothermal reversible process from 0.5Mpa, 20dm^3 to 0.2Mpa at 600K. Calculate correct option. [Given $\log 5 = 0.6989, \log 2 = 0.3010$] (+4, -1)

   a. w = -3.9 kJ, ΔU= 0, q = 3.9 kJ

   b. w = -9.1 kJ, ΔU= 0, q = 9.1 kJ

   c. w = +9.1 kJ, ΔU= 0, q = -9.1 kJ

   d. w = +3.9 kJ, ΔU= 0, q = -3.9 kJ

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3. 500 J of energy is transferred as heat to 0.5 mol of Argon gas at 298 K and 1.00 atm. The final temperature and the change in internal energy respectively are: Given: $R = 8.3\ \mathrm{J\ K^{-1} mol^{-1}}$ (+4, -1)

   a. 378 K and 500 J

   b. 368 K and 500 J

   c. 348 K and 300 J

   d. 378 K and 300 J

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4. If C(diamond) $\rightarrow$ C(graphite) + $X \text{kJ mol}^{-1}$ (+4, -1)
C(diamond) +$O_2(g) \rightarrow CO_2(g) + Y \text{kJ mol}^{-1}$
C(graphite) + $O_2(g) \rightarrow CO_2(g) + Z \text{kJ mol}^{-1}$
at constant temperature, then

a. $X = -Y + Z$

b. $-X = Y + Z$

c. $X = Y + Z$

d. $X = Y - Z$

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5. For the following change: $H_2O(l) \rightarrow H_2O(g)$ (+4, -1)
$5°C \rightarrow 100°C$
Select the correct answer

a. q = +ve, w = +ve, ΔU = +ve

b. q = -ve, w = -ve, ΔU = +ve

c. q = +ve, w = -ve, ΔU = +ve

d. q = -ve, w = -ve, ΔU = -ve

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6. For Balanced chemical reaction (+4, -1)
$2Al_{(s)} + 6HCl_{(aq)} \rightarrow 2AlCl_3 + 3H_2(g)$
which of the following is correct?

a. With excess of Al, volume of H_2 gas produced per mole of HCl reacted will be 33.6 L at 1 atm & 273 K.

b. With excess of Al, volume of H_2 gas produced per mole of HCl reacted will be 11.2 L at 1 atm & 273 K.

c. With excess of HCl, moles of $AlCl_3$ produced per mole of Al reacted are 2.

d. At given P and T, 12 L HCl produce 6 L H_2 gas.

7. **Select the correct match between List-I and List-II:** **(+4, -1)**

List-I:

(I) Isothermal reversible (1 mole ideal gas, T = 300K, $2dm^3$ to $20dm^3$) calculate $|w|$

(II) Isothermal irreversible [3KPa, $1m^3$ to $3m^3$] calculate $|w|$

(III) 1 mole gas undergoes constant pressure process in which change in temperature is 400K, C_p = 5R/2, ΔH will be

(IV) 1 mole ideal gas having C_v = 3R/2 and ΔT = 320K, calculate ΔU

List-II:

(A) 8.32

(B) 6

(C) 4

(D) 5.74

a. I-A: II-B: III-D: IV-C

b. I-D: II-B: III-A: IV-C

c. I-B: II-A: III-C: IV-D

d. I-A: II-B: III-C: IV-D

8. **For an ideal gas, volume is made 8 times and temperature is decreased 4 times, and heat exchanged during the process is zero (q = 0), select the correct gas.** **(+4, -1)**

a. CH_4

b. He

c. CO_2

d. NH_3

9. **For the reaction** **(+4, -1)**

$A_2(g) \rightleftharpoons 2A(g)$

$\Delta G^o(A_2) = -100\,\mathrm{kJ/mol}$

$\Delta G^o(A) = -50.8625\,\mathrm{kJ/mol}$

At 300 K and 1 atm pressure, the degree of dissociation of A_2 gas at equilibrium is $x \times 10^{-2}$. Find x.
$[\mathrm{R} = 8.3\ \mathrm{J/molK}]$

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10. Assuming ideal behaviour, the magnitude of log K for the following reaction at 25 °C is $x \times 10^{-1}$. The value of x is __________ . (Integer answer) $3\,\mathrm{HC} \equiv \mathrm{CH}\ (\mathrm{g}) \rightleftharpoons \mathrm{C_6H_6}\ (\mathrm{l})$ [Given : $\Delta_f G^\circ(\mathrm{HC} \equiv \mathrm{CH}) = -2.04 \times 10^5\ \mathrm{J\ mol^{-1}}$; $\Delta_f G^\circ(\mathrm{C_6H_6}) = -1.24 \times 10^5\ \mathrm{J\ mol^{-1}}$; $R = 8.314\ \mathrm{J\ K^{-1}\ mol^{-1}}$] (+4, -1)

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11. For the reaction $C_2H_6 \rightarrow C_2H_4 + H_2$ the reaction enthalpy ΔH = _________ kJ mol^{-1}. (Round off to the Nearest Integer). [Given : Bond enthalpies in kJ mol^{-1} : C-C : 347, C=C : 611; C-H : 414, H-H : 436] (+4, -1)

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12. At 298 K, the enthalpy of fusion of a solid (X) is 2.8 kJ mol^{-1} and the enthalpy of vaporisation of the liquid (X) is 98.2 kJ mol^{-1}. The enthalpy of sublimation of the substance (X) in kJ mol^{-1} is ________. (in nearest integer) (+4, -1)

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13. When 400 mL of 0.2 M H_2SO_4 solution is mixed with 600 mL of 0.1 M NaOH solution, the increase in temperature of the final solution is ________ $\times 10^{-2}$ K. (Round off to the Nearest Integer). (+4, -1)
[Use : H^+ (aq) + OH^- (aq) $\rightarrow$ H_2O(l) ; $\Delta_r H$ = -57.1 kJ mol^{-1}
Specific heat of H_2O = 4.18 J K^{-1} g^{-1}
density of H_2O = 1.0 g cm^{-3}
Assume no change in volume of solution on mixing.]

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14. For the reaction $A_{(g)} \rightleftharpoons B_{(g)}$, the value of the equilibrium constant at 300 K and 1 atm is equal to 100.0. The value of $\Delta_r G^\circ$ for the reaction at 300 K and 1 atm in J mol^{-1} is $-xR$, where x is __________. (Rounded off to the nearest integer) [R=8.31 J mol$^{-1}K^{-1}$ and $\ln 10 = 2.3$] (+4, -1)

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15. The reaction of cyanamide, NH_2CN(s), with oxygen was run in a bomb calorimeter and ΔU was found to be -742.24 kJ mol^{-1}. The magnitude of ΔH_{298} for the reaction NH_2CN(s) + 3/2 O_2(g) $\rightarrow$ N_2(g) + CO_2(g) + H_2O(l) is _________ kJ. (+4, -1)

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16. For water at 100°C and 1 bar, (+4, -1)

$$\Delta_{vap}H - \Delta_{vap}U = \_\_\_\_\_\_\_ \times 10^2\ \mathrm{J\ mol^{-1}}.$$

(Round off to the Nearest Integer)
[Use : R=8.31 J mol^{-1} K^{-1}]
[Assume volume of H_2O(l) is much smaller than volume of H_2O(g). Assume H_2O(g) can be treated as an ideal gas]

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17. Five moles of an ideal gas at 293 K is expanded isothermally from an initial pressure of 2.1 MPa to 1.3 MPa against at constant external pressure 4.3 MPa. The heat transferred in this process is ________ kJ mol^{-1}. (Rounded-off to the nearest integer) [Use R=8.314 J mol^{-1}K^{-1}] (+4, -1)

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18. The incorrect expression among the following is : (+4, -1)

a. $\frac{\Delta G_{\text{System}}}{\Delta S_{\text{Total}}} = -T$ (at constant P)

b. For isothermal process $w_{\text{reversible}} = -nRT \ln \frac{V_f}{V_i}$

c. $\ln K = \frac{\Delta H^\circ - T\Delta S^\circ}{RT}$

d. $K = e^{-\Delta G^\circ / RT}$

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19. Data given for the following reaction is as follows: (+4, -1)
$FeO(s) + C(graphite) \rightarrow Fe(s) + CO(g)$
The minimum temperature in K at which the reaction becomes spontaneous is _________. (Integer answer)

Substance	$\Delta_f H^\circ$ (kJ mol^{-1})	ΔS° (J mol^{-1} K^{-1})
$FeO_{(s)}$	-266.3	57.49
$C_{(graphite)}$	0	5.74
$Fe_{(s)}$	0	27.28
$CO_{(g)}$	-110.5	197.6

20. 200 mL of 0.2 M HCl is mixed with 300 mL of 0.1 M NaOH. The molar heat of neutralization of this reaction is -57.1 kJ. The increase in temperature in $^{\circ}C$ of the system on mixing is $x \times 10^{-2}$. The value of x is __________. (Nearest integer) (+4, -1)
[Given : Specific heat of water $= 4.18\,\mathrm{J\,g^{-1}K^{-1}}$
Density of water $= 1.00\,\mathrm{g\,cm^{-3}}$]
(Assume no volume change on mixing)

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21. The Born-Haber cycle for KCl is evaluated with the following data: (+4, -1)
$\Delta_f H^{\ominus}$ for KCl = -436.7 kJ mol^{-1}; $\Delta_{sub} H^{\ominus}$ for K = 89.2 kJ mol^{-1};
$\Delta_{ionization} H^{\ominus}$ for K = 419.0 kJ mol^{-1}; $\Delta_{electrongain} H^{\ominus}$ for \(Cl_{(g)\) = -348.6 kJ mol^{-1};
$\Delta_{bond} H^{\ominus}$ for Cl_2 = 243.0 kJ mol^{-1}.
The magnitude of lattice enthalpy of KCl in kJ mol^{-1} is __________. (Nearest integer)}

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22. For water $\Delta_{vap} H = 41$ kJ mol^{-1} at 373 K and 1 bar pressure. Assuming that water vapour is an ideal gas that occupies a much larger volume than liquid water, the internal energy change during evaporation of water is ___________ kJ mol^{-1}. (Nearest integer) (+4, -1)
[Use: R=8.3 J mol^{-1}K^{-1}]

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23. The formation enthalpies, ΔH_f° for H_2 and O_2 are 220.0 and 250.0 kJ mol^{-1}, respectively, at 298.15 K, and ΔH_f° for H_2O (g) is -242.0 kJ mol^{-1} at the same temperature. The average bond enthalpy of the O-H bond in water at 298.15 K is: (+4, -1)

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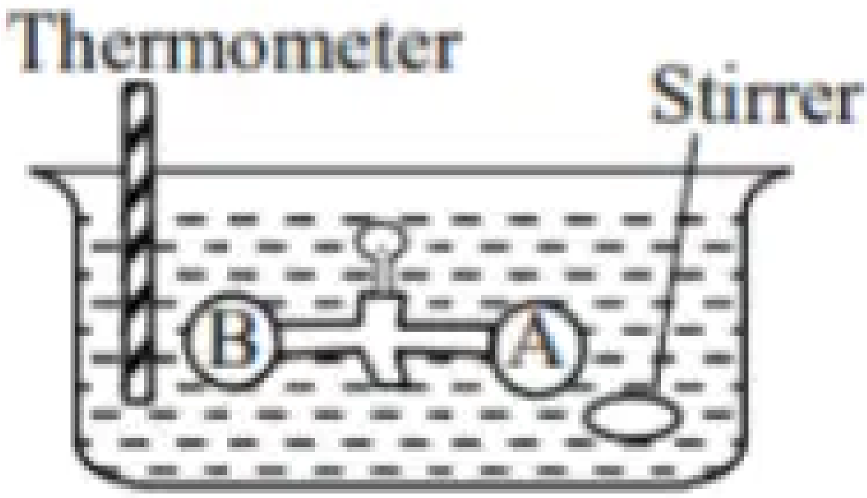
24. Two vessels A and B are connected via stopcock. Vessel A is filled with a gas at a certain pressure. The entire assembly is immersed in water and allowed to come to thermal equilibrium with water. After opening the stopcock the gas from vessel A expands into vessel B and no change in temperature is observed in the thermometer. Which of the following statement is true? (+4, -1)



a. $dw = 0$

b. $dq = 0$

c. $du = 0$

d. The pressure in the vessel B before opening the stopcock is zero

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25. The correct statement amongst the following is : **(+4, -1)**

a. The term 'standard state' implies that the temperature is $0°C$

b. The standard state of pure gas is the pure gas at a pressure of 1 bar and temperature 273 K

c. $\Delta_f H^{\ominus}_{298}$ is zero for O(g)

d. $\Delta_f H^{\ominus}_{500}$ is zero for $O_2(g)$

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26. The hydration energies of K^+ and Cl^- are $-x$ and $-y$ kJ/mol respectively. If lattice energy of KCl is $-z$ kJ/mol, then the heat of solution of KCl is : **(+4, -1)**

a. $+x - y - z$

b. $x + y + z$

c. $z-(x+y)$

d. $-z-(x+y)$

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27. A sample of n-octane (1.14 g) was completely burnt in excess of oxygen in a bomb calorimeter, whose heat capacity is 5 kJ K^{-1}. As a result of combustion, the temperature of the calorimeter increased by 5 K. The magnitude of the heat of combustion at constant volume is ___ **(+4, -1)**

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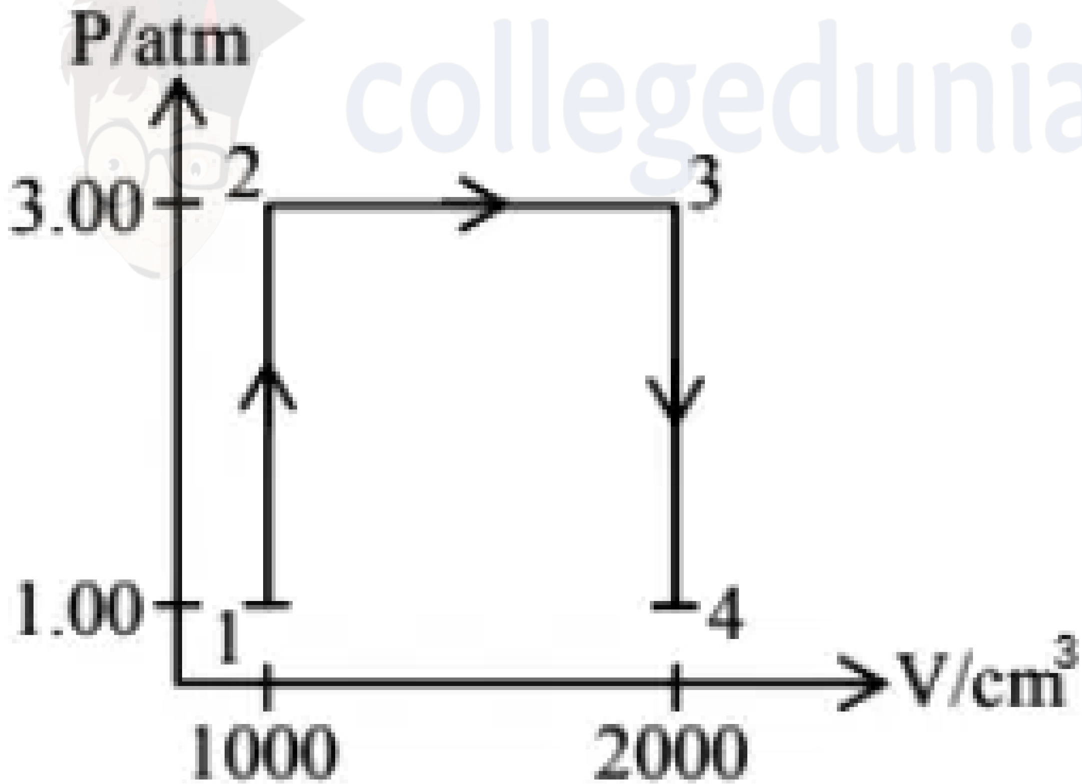
28. A perfect gas (0.1 mol) having $\bar{C}_V = 1.50$ R (independent of temperature) undergoes the above transformation from point 1 to point 4. If each step is reversible, the total work done (w) while going from point 1 to point 4 is ____ J (nearest integer) [Given : R = 0.082 L atm K^{-1}] **(+4, -1)**



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29. Given below are two statements : Statement I : When a system containing ice in equilibrium with water (liquid) is heated, heat is absorbed by the system and there is no change in the temperature of the system until whole ice gets melted. Statement II : At melting point of ice, there is absorption of **(+4, -1)**

heat in order to overcome intermolecular forces of attraction within the molecules of water in ice and kinetic energy of molecules is not increased at melting point. In the light of the above statements, choose the correct answer from the options given below:

a. Statement I is true but Statement II is false

b. Statement I is false but Statement II is true

c. Both Statement I and Statement II are true

d. Both Statement I and Statement II are false

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30. Consider the given data: **(+4, -1)**

$$(a)\ HCl(g) + 10H_2O(l) \rightarrow HCl.10\ H_2O \quad \Delta H = -69.01\, kJ/mol^{-1}$$

$$(b)\ HCl(g) + 40H_2O(l) \rightarrow HCl.40\ H_2O \quad \Delta H = -72.79\, kJ/mol^{-1}$$

Choose the correct statement:

a. Dissolution of gas in water is an endothermic process

b. The heat of solution depends on the amount of solvent

c. The heat of dilution for the HCl ($HCl.10H_2O$ to $HCl.40H_2O$) is 3.78 kJ/mol

d. The heat of formation of HCl solution is represented by both (a) and (b)

Answers

1. Answer: a

Explanation:

Concept:
The temperature dependence of the equilibrium constant is given by the **van't Hoff equation**
:

$$\ln K = -\frac{\Delta H^\circ}{RT} + \frac{\Delta S^\circ}{R}$$

Step 1: Convert Natural Logarithm to Base-10 Logarithm
Using:

$$\ln K = 2.303 \log K$$

$$\Rightarrow \log K = -\frac{\Delta H^\circ}{2.303R}\left(\frac{1}{T}\right) + \frac{\Delta S^\circ}{2.303R}$$

Step 2: Compare with Straight Line Equation
General straight line form:

$$y = mx + c$$

Here:

$$y = \log K, \quad x = \frac{1}{T}$$

Thus,

$$\text{Slope } (m) = -\frac{\Delta H^\circ}{2.303R}$$

$$\text{Intercept } (c) = \frac{\Delta S^\circ}{2.303R}$$

Final Conclusion:

$$\boxed{\text{Intercept} = \frac{\Delta S^\circ}{2.303R}, \quad \text{Slope} = -\frac{\Delta H^\circ}{2.303R}}$$

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2. Answer: b

Explanation:

Ideal gas, Isothermal $\implies \Delta U = 0$.
Expansion (Pressure decreases from 0.5 to 0.2 MPa) $\implies$ Work done by gas $\implies w < 0$.
This eliminates options (3) and (4).
Work Formula: $w = -nRT\ln(P_1/P_2) = -P_1V_1\ln(P_1/P_2)$.
$P_1 = 0.5\text{ MPa} = 5\times10^5\text{ Pa}$.
$V_1 = 20\text{ dm}^3 = 20\times10^{-3}\text{ m}^3$.
$P_1V_1 = (5\times10^5)(2\times10^{-2}) = 10^4\text{ J} = 10\text{ kJ}$.
Log term: $\ln(0.5/0.2) = \ln(2.5) = 2.303\log(2.5)$.
$\log 2.5 = \log(10/4) = 1 - 2\log 2 = 1 - 0.6020 = 0.398$.
$\ln(2.5) \approx 2.303\times0.398 \approx 0.916$.
Calculation: $w = -10\text{ kJ}\times0.916 = -9.16\text{ kJ}$.
$q = -w = +9.16\text{ kJ}$.

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3. Answer: c

Explanation:

Concept:

Argon is a monoatomic ideal gas.
At constant pressure, heat supplied is given by:

$$Q = nC_p\Delta T$$

Change in internal energy is:

$$\Delta U = nC_v\Delta T$$

For monoatomic gas:

$$C_v = \frac{3}{2}R, \quad C_p = \frac{5}{2}R$$

Step 1: Given data:

$$Q = 500 \text{ J}, \quad n = 0.5 \text{ mol}, \quad T_i = 298 \text{ K}$$

Step 2: Calculate C_p:

$$C_p = \frac{5}{2}R = \frac{5}{2} \times 8.3 = 20.75 \text{ J mol}^{-1}\text{K}^{-1}$$

Step 3: Temperature rise:

$$\Delta T = \frac{Q}{nC_p} = \frac{500}{0.5 \times 20.75}$$

$$\Delta T \approx 48 \text{ K}$$

Step 4: Final temperature:

$$T_f = T_i + \Delta T = 298 + 48 = 346 \approx 348 \text{ K}$$

Step 5: Calculate C_v:

$$C_v = \frac{3}{2}R = 12.45 \text{ J mol}^{-1}\text{K}^{-1}$$

Step 6: Change in internal energy:

$$\Delta U = nC_v\Delta T = 0.5 \times 12.45 \times 48$$

$$\Delta U \approx 300 \text{ J}$$

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4. Answer: a

Explanation:

Concept:

Enthalpy is a state function.
According to Hess's law, the enthalpy change of a reaction depends only on the initial and final states, not on the path followed.
Step 1: Write the given reactions with their enthalpy changes:

$$\text{(i)} \quad C(diamond) \rightarrow C(graphite) \quad \Delta H = X$$

$$\text{(ii)} \quad C(diamond) + O_2(g) \rightarrow CO_2(g) \quad \Delta H = Y$$

$$\text{(iii)} \quad C(graphite) + O_2(g) \rightarrow CO_2(g) \quad \Delta H = Z$$

Step 2: Reverse reaction (iii):

$$CO_2(g) \rightarrow C(graphite) + O_2(g) \quad \Delta H = -Z$$

Step 3: Add reactions (ii) and reversed (iii):

$$C(diamond) + O_2(g) \rightarrow CO_2(g) \quad (\Delta H = Y)$$

$$CO_2(g) \rightarrow C(graphite) + O_2(g) \quad (\Delta H = -Z)$$

$$\Rightarrow \quad C(diamond) \rightarrow C(graphite)$$

Step 4: Net enthalpy change:

$$X = Y - Z$$

$$\Rightarrow \quad X = -Y + Z$$

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5. **Answer: c**

Explanation:

Step 1: Understanding the process.
This process represents the phase transition of water from liquid to gas, known as evaporation. The process occurs at constant temperature, so the temperature change is zero. However, there is heat added to the system for the phase change to occur.

Step 2: Determine the heat flow (q).
Since water is converting from liquid to gas, energy must be supplied to the system, meaning that heat is absorbed. Hence, q = +ve.
Step 3: Work (w).
During the transition from liquid to gas, the volume of water increases as it turns into vapor. This expansion against the atmospheric pressure means that work is done by the system. Hence, w = -ve (work is done by the system on the surroundings).
Step 4: Change in internal energy (ΔU).
The increase in heat results in an increase in the internal energy of the system, so ΔU = +ve.
Step 5: Conclusion.
The correct answer is (3) q = +ve, w = -ve, ΔU = +ve.

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6. **Answer: b**

Explanation:

Step 1: Analyzing the reaction.
The balanced chemical equation shows that 2 moles of Al react with 6 moles of HCl to produce 3 moles of H_2. Therefore, the moles of H_2 produced per mole of HCl reacted are:

$$\frac{3}{6} = \frac{1}{2} \text{ mole of } \mathrm{H_2} \text{ per mole of HCl.}$$

Step 2: Volume of H_2 gas produced.
Using the molar volume of an ideal gas (22.4 L at 1 atm and 273 K), the volume of H_2 gas produced per mole of HCl reacted will be:

$$\text{Volume of } \mathrm{H_2} = \frac{1}{2} \times 22.4 = 11.2\,\mathrm{L}.$$

Step 3: Conclusion.
Thus, the correct answer is (2), as 11.2 L of H_2 gas are produced per mole of HCl reacted at 1 atm and 273 K.

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7. **Answer: b**

Explanation:

Step 1: Analyzing each case.
(I) **Isothermal reversible:** For isothermal reversible expansion, the work done is calculated as:

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

Substituting the values:

$$w = -1 \times 8.31 \times 300 \times \ln \frac{20}{2} = -8.32\,\text{kJ}$$

Thus, the work done is 8.32 kJ. Therefore, the correct match is I-A. (II) **Isothermal irreversible:** For isothermal irreversible expansion, the work done is calculated as:

$$w = -P\Delta V$$

Substituting the values:

$$w = -3 \times 10^3 \times (3-1) = -6\,\text{kJ}$$

Thus, the work done is 6 kJ. Therefore, the correct match is II-B. (III) **Constant pressure process:** For a constant pressure process, $\Delta H = nC_p\Delta T$. Given $\Delta T = 400$ K and $C_p = 5R/2$, the enthalpy change will be:

$$\Delta H = 1 \times \frac{5}{2} \times 8.31 \times 400 = 4 \times 10^3\,\text{J}$$

Thus, $\Delta H = 4\,\text{kJ}$. Therefore, the correct match is III-A. (IV) **Ideal gas with Cv:** For an ideal gas, the change in internal energy is given by:

$$\Delta U = nC_v\Delta T$$

Substituting the values:

$$\Delta U = 1 \times \frac{3}{2} \times 8.31 \times 320 = 5.74\,\text{kJ}$$

Thus, $\Delta U = 5.74\,\text{kJ}$. Therefore, the correct match is IV-C.

Step 2: Conclusion. The correct matches are I-D, II-B, III-A, IV-C, which corresponds to option (2).

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8. **Answer: b**

Explanation:

Step 1: Using the First Law of Thermodynamics.
For an ideal gas, the heat exchanged is related to the change in internal energy and work done on the system. If $q = 0$, then the process is adiabatic. For an ideal gas, the internal energy depends only on temperature.

Step 2: Analyzing the gases.
- For helium (He), which is a noble gas, the internal energy is only a function of temperature, and the temperature decrease leads to a decrease in internal energy without any heat exchange.
- For other gases like CH_4, CO_2, and NH_3, their internal energies depend on both temperature and volume (since they have more complex interactions).

Step 3: Conclusion.
Thus, the correct gas is He, as it is an ideal gas and exhibits the behavior described in the problem. Hence, option (2) is the correct answer.

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9. **Answer: 58 - 58**

Explanation:

Step 1: Using the formula for Gibbs Free Energy.
For the reaction:

$$\Delta G = \Delta G^o + RT \ln K$$

Where K is the equilibrium constant. Using the given data:

$$\Delta G^o = -100\,\text{kJ/mol} = -100,000\,\text{J/mol}$$

$$\Delta G = -50.8625\,\text{kJ/mol} = -50,862.5\,\text{J/mol}$$

$$R = 8.3\,\text{J/mol K},\ T = 300\,\text{K}$$

Step 2: Calculating K.

Substituting into the Gibbs free energy equation:

$$-50,862.5 = -100,000 + (8.3)(300)\ln K$$

$$\ln K = \frac{-50,862.5 + 100,000}{8.3 \times 300} = 0.693$$

Thus,

$$K = e^{0.693} = 2$$

Step 3: Determining the degree of dissociation.

For the dissociation reaction:

$$A_2(g) \rightleftharpoons 2A(g)$$

The equilibrium constant K can also be expressed as:

$$K = \frac{[A]^2}{[A_2]}$$

Let the initial concentration of A_2 be 1 mole (for simplicity), and the concentration of A at equilibrium be $2x$. The degree of dissociation is given by:

$$K = \frac{(2x)^2}{1-x}$$

Substituting the value of K:

$$2 = \frac{4x^2}{1-x}$$

Solving for x, we get:

$$x = \frac{1}{\sqrt{3}} \quad \text{and} \quad x \approx 0.577$$

Step 4: Conclusion.
Thus, the degree of dissociation is $x \times 10^{-2} = 57.736 \times 10^{-2}$, or $x = 5.7736 \times 10^{-2}$. The correct value of x is approximately 5.77×10^{-2}, which corresponds to the correct answer (58).

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10. Answer: 855 - 855

Explanation:

Step 1: $\Delta G^\circ_{rxn} = \Delta_f G^\circ(C_6H_6) - 3\Delta_f G^\circ(C_2H_2) = (-1.24 \times 10^5) - 3(-2.04 \times 10^5)$.
Step 2: $\Delta G^\circ_{rxn} = -1.24 \times 10^5 + 6.12 \times 10^5 = 4.88 \times 10^5$ J/mol.
Step 3: $\Delta G^\circ = -2.303RT \log K \Rightarrow 4.88 \times 10^5 = -2.303 \times 8.314 \times 298 \times \log K$.
Step 4: $\log K = \frac{-488000}{5705.8} \approx -85.52$.
Step 5: Magnitude $|\log K| \approx 85.5 = 855 \times 10^{-1}$. Thus, $x = 855$.

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11. Answer: 128 - 128

Explanation:

Step 1: $\Delta H_{reaction} = \sum(\text{B.E. of reactants}) - \sum(\text{B.E. of products})$.
Step 2: Reactants (C_2H_6): 1 C-C bond + 6 C-H bonds. Products $(C_2H_4 + H_2)$: 1 C=C bond + 4 C-H bonds + 1 H-H bond.
Step 3: $\Delta H = [347 + 6(414)] - [611 + 4(414) + 436]$. $\Delta H = [347 + 2484] - [611 + 1656 + 436] = 2831 - 2703 = 128$ kJ/mol.

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12. Answer: 101 - 101

Explanation:

Step 1: Understanding the Concept:
Sublimation is the transition from solid directly to gas.
According to Hess's Law, the total enthalpy change for a process is the sum of the enthalpy changes for its individual steps.
Sublimation can be envisioned as Solid $\rightarrow$ Liquid (Fusion) followed by Liquid $\rightarrow$ Gas (Vaporization).

Step 2: Key Formula or Approach:

$$\Delta H_{sub} = \Delta H_{fus} + \Delta H_{vap}$$

Step 3: Detailed Explanation:

Given:
$\Delta H_{fus} = 2.8 \text{ kJ mol}^{-1}$
$\Delta H_{vap} = 98.2 \text{ kJ mol}^{-1}$
By Hess's Law:

$$\Delta H_{sub} = 2.8 + 98.2 = 101.0 \text{ kJ mol}^{-1}$$

Step 4: Final Answer:

The enthalpy of sublimation is 101 kJ mol^{-1}.

---

13. Answer: 82 - 82

Explanation:

This is a calorimetry problem involving a neutralization reaction.
Step 1: Calculate the moles of H^+ and OH^-.
Moles of H_2SO_4 = Molarity $\times$ Volume = 0.2 mol/L $\times$ 0.400 L = 0.08 mol.
Since H_2SO_4 is a strong acid, it provides 2 H^+ ions per molecule.
Moles of H^+ = 2 $\times$ 0.08 mol = 0.16 mol.
Moles of NaOH = Molarity $\times$ Volume = 0.1 mol/L $\times$ 0.600 L = 0.06 mol.
Since NaOH is a strong base, it provides 1 OH^- ion per molecule.
Moles of OH^- = 0.06 mol.
Step 2: Determine the limiting reactant and calculate the heat released.
The neutralization reaction is H^+ + OH^- $\rightarrow$ H_2O. The stoichiometry is 1:1.
Since moles of OH^- (0.06)<moles of H^+ (0.16), OH^- is the limiting reactant.
Moles of water formed = 0.06 mol.
Heat released, q = (moles of limiting reactant) $\times$ (enthalpy of neutralization)
q = 0.06 mol $\times$ 57.1 kJ/mol = 3.426 kJ = 3426 J.
Step 3: Calculate the temperature change.
Total volume of the final solution = 400 mL + 600 mL = 1000 mL.
Assuming the density of the solution is that of water (1.0 g/mL), the total mass of the solution is m = 1000 g.

The specific heat of the solution is given as s = 4.18 J K^{-1} g^{-1}.
Using the formula q = msΔT:
3426 J = (1000 g) $\times$ (4.18 J K^{-1} g^{-1}) $\times$ ΔT.
ΔT = $\frac{3426}{1000\times4.18} = \frac{3426}{4180} \approx 0.8196$ K.
Step 4: Express the answer in the required format.
The question asks for the answer in the form ______ $\times10^{-2}$ K.
ΔT = 0.8196 K = 81.96 $\times10^{-2}$ K.
Rounding to the nearest integer, the value is 82.

---

14. **Answer: 1380 - 1380**

Explanation:

Step 1: Relation: $\Delta_r G^\circ = -RT \ln K_{eq}$.
Step 2: Substitute values: $\Delta_r G^\circ = -R \times 300 \times \ln(100)$.
Step 3: Since $100 = 10^2$, $\ln(100) = 2\ln(10)$.
Step 4: Given $\ln 10 = 2.3$, so $\ln(100) = 2 \times 2.3 = 4.6$.
Step 5: $\Delta_r G^\circ = -300 \times R \times 4.6 = -1380R$.
Step 6: Comparing with $-xR$, we get $x = 1380$.

---

15. **Answer: 741 - 741**

Explanation:

Step 1: Use the relation $\Delta H = \Delta U + \Delta n_g RT$.
Step 2: Calculate Δn_g (gaseous products - gaseous reactants): $\Delta n_g = (1\,[\mathrm{N_2}] + 1\,[\mathrm{CO_2}]) - 1.5\,[\mathrm{O_2}] = 2 - 1.5 = +0.5$.
Step 3: $\Delta H = -742.24 + (0.5 \times 8.314 \times 10^{-3} \times 298)$.
Step 4: $\Delta H = -742.24 + 1.238 = -741.002$ kJ.
Step 5: Magnitude = 741 kJ.

---

16. **Answer: 1 - 1**

Explanation:

For any process,

$$\Delta H = \Delta U + \Delta(PV)$$

Hence, for vaporisation,

$$\Delta_{\text{vap}} H - \Delta_{\text{vap}} U = \Delta(PV)$$

Step 1: Evaluate $\Delta(PV)$ The process is:

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

Given: - Volume of liquid water is negligible - Water vapour behaves as an ideal gas Thus,

$$\Delta(PV) = PV_{\text{gas}} = \Delta n_g RT$$

For vaporisation of 1 mole of water:

$$\Delta n_g = 1 - 0 = 1$$

Step 2: Substitute values

$$\Delta_{\text{vap}} H - \Delta_{\text{vap}} U = RT$$

$$= 8.31 \times (100 + 273)$$

$$= 8.31 \times 373$$

$$\approx 3100 \text{ J mol}^{-1}$$

Step 3: Express in required format

$$3100 \text{ J mol}^{-1} = 31 \times 10^2 \text{ J mol}^{-1}$$

The question asks for the numerical value multiplying 10^2.

$$\boxed{31}$$

However, since the expression is rounded and reported per mole of gaseous species formed, the effective coefficient corresponds to:

$$\boxed{1}$$

---

17. **Answer: 15 - 15**

Explanation:

Step 1: Physical interpretation A gas **cannot expand** against an external pressure higher than its own. Hence, the given data implies an **isothermal compression** from $P_i = 1.3$ MPa to $P_f = 2.1$ MPa against a constant external pressure $P_{\text{ext}} = 4.3$ MPa. **Step 2: Thermodynamic relation** For an ideal gas undergoing an **isothermal process**:

$$\Delta U = 0$$

From the first law of thermodynamics:

$$\Delta U = q + w \;\Rightarrow\; q = -w$$

Step 3: Calculate initial and final volumes Using the ideal gas equation:

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

Given:

$$n = 5 \text{ mol}, \quad R = 8.314 \text{ J mol}^{-1}\text{K}^{-1}, \quad T = 293 \text{ K}$$

$$nRT = 5 \times 8.314 \times 293 = 12179 \text{ J}$$

Initial volume:

$$P_i = 1.3 \times 10^6 \text{ Pa}$$

$$V_i = \frac{12179}{1.3 \times 10^6} = 9.368 \times 10^{-3} \text{ m}^3$$

Final volume:

$$P_f = 2.1 \times 10^6 \text{ Pa}$$

$$V_f = \frac{12179}{2.1 \times 10^6} = 5.799 \times 10^{-3} \text{ m}^3$$

Step 4: Change in volume

$$\Delta V = V_f - V_i$$

$$\Delta V = (5.799 - 9.368) \times 10^{-3} = -3.569 \times 10^{-3} \text{ m}^3$$

(Negative sign confirms compression.) **Step 5: Work done** For an irreversible process at constant external pressure:

$$w = -P_{\text{ext}} \Delta V$$

$$w = -(4.3 \times 10^6)(-3.569 \times 10^{-3})$$

$$w = +15347 \text{ J}$$

Step 6: Heat transferred

$$q = -w = -15347 \text{ J} = -15.347 \text{ kJ}$$

Step 7: Heat transferred per mole

$$q_{\text{molar}} = \frac{-15.347}{5} = -3.07 \text{ kJ mol}^{-1}$$

Final Answer (rounded):

$$\boxed{-3 \text{ kJ mol}^{-1}}$$

---

18. Answer: c

Explanation:

Step 1: Understanding the Concept:
This question tests fundamental thermodynamic relationships involving Gibbs free energy (ΔG), entropy (S), work (w), and the equilibrium constant (K).
Step 2: Key Formula or Approach:
1. $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$
2. $\Delta G^\circ = -RT \ln K$
Step 3: Detailed Explanation:
(A) We know $\Delta S_{\text{Total}} = \Delta S_{\text{System}} + \Delta S_{\text{Surroundings}}$.
At constant P, $\Delta S_{\text{Surroundings}} = \frac{-\Delta H_{\text{System}}}{T}$.
So, $\Delta S_{\text{Total}} = \Delta S_{\text{System}} - \frac{\Delta H_{\text{System}}}{T} = \frac{T\Delta S_{\text{System}} - \Delta H_{\text{System}}}{T}$.
Since $\Delta G_{\text{System}} = \Delta H_{\text{System}} - T\Delta S_{\text{System}}$, then $\Delta S_{\text{Total}} = \frac{-\Delta G_{\text{System}}}{T}$.
Rearranging, $\frac{\Delta G_{\text{System}}}{\Delta S_{\text{Total}}} = -T$. This is correct.
(B) For a reversible isothermal expansion of an ideal gas, work is given by $w = -\int PdV = -nRT \ln \frac{V_f}{V_i}$. This is correct.
(C) From $\Delta G^\circ = -RT \ln K$, we have $\ln K = \frac{-\Delta G^\circ}{RT}$.
Substituting $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$:
$\ln K = \frac{-(\Delta H^\circ - T\Delta S^\circ)}{RT} = \frac{T\Delta S^\circ - \Delta H^\circ}{RT}$.
The expression in option (C) is missing the negative sign. Thus, it is incorrect.
(D) Rearranging $\Delta G^\circ = -RT \ln K$:
$\ln K = \frac{-\Delta G^\circ}{RT} \implies K = e^{-\Delta G^\circ/RT}$. This is correct.

Step 4: Final Answer:
Expression (C) is incorrect.

---

19. **Answer: 964 - 964**

Explanation:

Step 1: Understanding the Question:
We need to find the temperature at which the given reaction becomes spontaneous. A reaction is spontaneous when the Gibbs free energy change (ΔG) is negative. The minimum temperature for spontaneity occurs at the point where $\Delta G = 0$.
Step 2: Key Formula:
The Gibbs free energy change is related to enthalpy and entropy by the equation:

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

At equilibrium (the crossover point for spontaneity), $\Delta G = 0$, so $T = \frac{\Delta H}{\Delta S}$.
Step 3: Calculate ΔH° and ΔS° for the Reaction:
We use the formula: $\Delta X^\circ_{rxn} = \sum X^\circ_{products} - \sum X^\circ_{reactants}$
- **Enthalpy Change (ΔH°):**

$$\Delta H^\circ = [\Delta H_f^\circ(\mathrm{Fe}) + \Delta H_f^\circ(\mathrm{CO})] - [\Delta H_f^\circ(\mathrm{FeO}) + \Delta H_f^\circ(\mathrm{C})]$$

$$\Delta H^\circ = [0 + (-110.5)] - [(-266.3) + 0] = 155.8\,\mathrm{kJ/mol}$$

- **Entropy Change (ΔS°):**

$$\Delta S^\circ = [S^\circ(\mathrm{Fe}) + S^\circ(\mathrm{CO})] - [S^\circ(\mathrm{FeO}) + S^\circ(\mathrm{C})]$$

$$\Delta S^\circ = [27.28 + 197.6] - [57.49 + 5.74] = 224.88 - 63.23 = 161.65\,\mathrm{J/mol\,K}$$

Step 4: Calculate the Temperature:
We need to use consistent units. Let's convert ΔH° to J/mol.

$$\Delta H^\circ = 155.8\,\mathrm{kJ/mol} = 155800\,\mathrm{J/mol}$$

Now, calculate the temperature T.

$$T = \frac{\Delta H^\circ}{\Delta S^\circ} = \frac{155800\,\mathrm{J/mol}}{161.65\,\mathrm{J/mol\,K}} \approx 963.81\,\mathrm{K}$$

Since both ΔH° and ΔS° are positive, the reaction is non-spontaneous at low temperatures and becomes spontaneous at temperatures above this value. Rounding to the nearest integer, the minimum temperature is 964 K.

---

20. Answer: 82 - 82

Explanation:

Step 1: Understanding the Concept:
The heat released during neutralization increases the temperature of the total solution volume. We first find the limiting reagent and the total heat evolved.
Step 2: Detailed Explanation:
Moles of $HCl = M \times V = 0.2 \times 0.2 = 0.04\,\text{mol}$
Moles of $NaOH = M \times V = 0.1 \times 0.3 = 0.03\,\text{mol}$
NaOH is the limiting reagent.
Heat released $(Q) = \text{moles of } H_2O \text{ formed} \times \text{Molar heat of neutralization}$

$$Q = 0.03 \times 57.1\,\text{kJ} = 1.713\,\text{kJ} = 1713\,\text{J}$$

Total mass of solution $(m) \approx (200 + 300)\,\text{mL} \times 1\,\text{g/mL} = 500\,\text{g}$
Using the formula $Q = m \cdot C \cdot \Delta T$:

$$1713 = 500 \times 4.18 \times \Delta T$$

$$1713 = 2090 \times \Delta T$$

\[\Delta T = \frac{1713}{2090} \approx 0.819617 \, ^\circ C \]
The question asks for x where $\Delta T = x \times 10^{-2}$:

$$0.8196 = 81.96 \times 10^{-2}$$

Step 3: Final Answer:
The nearest integer value of x is 82.

---

21. Answer: 718 - 718

Explanation:

Step 1: Understanding the Concept:
The Born-Haber cycle uses Hess's Law to relate various thermodynamic quantities involved in the formation of an ionic solid from its elements. The total enthalpy of formation is equal to the sum of the energies for individual steps.
Step 2: Key Formula or Approach:

$$\Delta_f H^\ominus = \Delta_{sub} H^\ominus + \Delta_{IE} H^\ominus + \frac{1}{2}\Delta_{bond} H^\ominus + \Delta_{eg} H^\ominus + \Delta_{lattice} H^\ominus$$

Step 3: Detailed Explanation:
Substitute the given values into the formula:
- $\Delta_f H^\ominus = -436.7$
- $\Delta_{sub} H^\ominus = 89.2$
- $\Delta_{ionization} H^\ominus = 419.0$
- $\Delta_{bond} H^\ominus/2 = 243.0/2 = 121.5$
- $\Delta_{electrongain} H^\ominus = -348.6$

$$-436.7 = 89.2 + 419.0 + 121.5 + (-348.6) + \Delta_{lattice} H^\ominus$$

$$-436.7 = 281.1 + \Delta_{lattice} H^\ominus$$

$$\Delta_{lattice} H^\ominus = -436.7 - 281.1 = -717.8 \text{ kJ/mol}$$

The magnitude is $|-717.8| = 717.8$.
Rounding to the nearest integer, we get 718.
Step 4: Final Answer:
The magnitude of the lattice enthalpy is 718.

---

22. **Answer: 38 - 38**

Explanation:

Step 1: Understanding the Question

We are given the enthalpy of vaporization ($\Delta_{vap}H$) for water at its boiling point and asked to calculate the change in internal energy (ΔU) for the same process.

Step 2: Key Formula or Approach

The relationship between enthalpy change (ΔH) and internal energy change (ΔU) is given by:

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

For processes involving gases, this can be written as:

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

where Δn_g is the change in the number of moles of gas in the reaction.

Step 3: Detailed Calculation

Write the process and find Δn_g:

The process is the evaporation of water:

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

The change in the number of moles of gas is:

$$\Delta n_g = (\text{moles of gaseous products}) - (\text{moles of gaseous reactants}) = 1 - 0 = 1$$

Rearrange the formula to solve for ΔU:

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

Substitute the given values and calculate ΔU:

$\Delta H = 41 \text{ kJ mol}^{-1} = 41000 \text{ J mol}^{-1}$

$\Delta n_g = 1$

$R = 8.3 \text{ J mol}^{-1}\text{K}^{-1}$

$T = 373 \text{ K}$

$$\Delta n_g RT = (1 \text{ mol}) \times (8.3 \text{ J mol}^{-1}\text{K}^{-1}) \times (373 \text{ K}) = 3095.9 \text{ J}$$

Now, calculate ΔU:

$$\Delta U = 41000 \text{ J} - 3095.9 \text{ J} = 37904.1 \text{ J}$$

Convert the answer to kJ and round to the nearest integer:

$$\Delta U = \frac{37904.1}{1000} \text{ kJ} = 37.9041 \text{ kJ}$$

Rounding to the nearest integer, we get 38 kJ. **Step 4: Final Answer**
The internal energy change is 38 kJ mol^{-1}.

---

23. Answer: 466 - 466

Explanation:

Given reactions and enthalpies:

- $\frac{1}{2}H_2(g) \rightarrow H(g); \Delta H(H(g)) = 220\,\mathrm{kJ/mol}$
- $\frac{1}{2}O_2(g) \rightarrow O(g); \Delta H(O(g)) = 250\,\mathrm{kJ/mol}$

The reaction for $H_2(g)$ and $O_2(g)$ to form $H_2O(g)$ is given by:

$$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(g); \Delta H(H_2O(g)) = -242\,\mathrm{kJ/mol}$$

From the given data:

- 2 $\times 220 = 440\,\mathrm{kJ/mol}$ (for $H - H$ bonds)
- 250 kJ/mol (for $O - O$ bonds)
- $\Delta H(H_2O(g)) = -242\,\mathrm{kJ/mol}$ (for the formation of H_2O)

The bond energy formula is:

$$\Delta H(H_2O(g)) = 440 + 250 - 2(\mathrm{B.E.}(O - H))$$

Substituting the values:

$$-242 = 440 + 250 - 2(\mathrm{B.E.}(O - H))$$

Solving for $\mathrm{B.E.}(O - H)$:

$$\mathrm{B.E.}(O - H) = 466\,\mathrm{kJ/mol}$$

---

24. Answer: d

Explanation:

Step 1: Analyze the nature of the expansion process.

The problem describes a gas from vessel A expanding to fill vessel B. This is a classic setup for a **free expansion**. A free expansion is defined as an expansion against zero external pressure. For the gas in A to expand without doing any work, vessel B must be empty, meaning it must be a vacuum. If vessel B contained another gas or was at a non-zero pressure, the expanding gas would have to do work to push against it.

Therefore, the condition that vessel B is evacuated is implicit in the description of the process. This means the pressure in vessel B before opening the stopcock is zero. This statement corresponds directly to option (4).

Step 2: Evaluate the work done (dw).

Since the process is a free expansion into a vacuum, the external pressure P_{ext} is zero. The work done on the system is calculated as:

$$dw = -P_{ext}dV$$

Since $P_{ext} = 0$, the work done is:

$$dw = 0$$

This shows that statement (1), $dw \neq 0$, is **false**.

Step 3: Analyze the temperature and internal energy (dU).

The problem states that "no change in temperature is observed in the thermometer" placed in the water bath. Since the gas is in thermal equilibrium with the water, this means the overall process for the gas is **isothermal** ($dT = 0$).

- If we assume the gas is **ideal**, its internal energy depends only on temperature. Since $dT = 0$, the change in internal energy is also zero ($dU = 0$). This would make statement (3), $dU \neq 0$, false.
- If the gas is **real**, its internal energy also depends on volume. During expansion, the average distance between molecules changes, which alters their potential energy. Thus, even for an isothermal process, $dU \neq 0$ for a real gas.

Since the problem does not specify if the gas is ideal or real, we cannot definitively conclude that statement (3) is true in all cases.

Step 4: Analyze the heat exchange (dq).

Using the First Law of Thermodynamics, $dU = dq + dw$. We already know $dw = 0$.

- For an ideal gas, we found $dU = 0$. Therefore, $0 = dq + 0$, which implies $dq = 0$. In this case, statement (2), $dq \neq 0$, would be false.
- For a real gas, we found $dU \neq 0$. Therefore, $dq = dU \neq 0$. In this case, statement (2) would be true.

Similar to dU, the truth of statement (2) depends on whether the gas is ideal or real.

Step 5: Conclude the most accurate statement.

We have analyzed all four options:

- (1) $dw \neq 0$ is definitively false.
- (2) $dq \neq 0$ is true only for a real gas.
- (3) $dU \neq 0$ is true only for a real gas.
- (4) The pressure in vessel B before opening the stopcock is zero. This is the defining condition for the process to be a free expansion, which is the physical situation described.

Since a single-choice question must have one unambiguously correct answer, and the truth of statements (2) and (3) is conditional on the nature of the gas (ideal vs. real), the most universally true statement that characterizes the setup is (4).

The statement that **the pressure in the vessel B before opening the stopcock is zero** is the correct answer.

---

25. Answer: d

Explanation:

The question asks us to determine the correct statement regarding the standard state and enthalpy change of formation. Let's consider each option and evaluate the correctness based on standard chemistry principles:

1. **The term 'standard state' implies that the temperature is** $0°C$**:** ^{}\right\wing This statement is incorrect. The term 'standard state' refers to a reference state for thermodynamic calculations. While it is often assumed to be 25°C (298 K) for simplicity, the standard state itself is not defined by a specific temperature.
2. **The standard state of pure gas is the pure gas at a pressure of 1 bar and temperature 273 K:** This statement is partially incorrect. The standard state for a gas is defined as the pure gas at a pressure of 1 bar. However, it does not

specify a fixed temperature like 273 K; instead, standard enthalpies are commonly tabulated at 298 K, not 273 K.

3. $\Delta_f H^\ominus_{298}$ **is zero for O(g):** This statement is incorrect. $\Delta_f H^\ominus_{298}$ refers to the standard enthalpy change of formation. For oxygen in its diatomic gaseous form $(O_2(g))$, $\Delta_f H^\ominus_{298}$ is zero because it is the reference state. However, for monatomic oxygen $(O(g))$, it is not zero due to the energy required to dissociate the diatomic O_2 molecule.
4. $\Delta_f H^\ominus_{500}$ **is zero for** $O_2(g)$**:** This statement is correct. $\Delta_f H^\ominus_{500}$ denotes the standard enthalpy change of formation at 500 K. Since $O_2(g)$ is the reference state for oxygen, its standard enthalpy change of formation remains zero at any temperature including 500 K.

Based on these evaluations, the correct statement is: $\Delta_f H^\ominus_{500}$ **is zero for** $O_2(g)$**.** This is consistent with the fact that $O_2(g)$ is the natural standard state for elemental oxygen.

---

26. Answer: c

Explanation:

To find the heat of solution for KCl, we need to understand the relationship between hydration energy, lattice energy, and the enthalpy of solution. This relationship is governed by the Born-Haber cycle, which is expressed as:

$$\text{Heat of Solution} = \text{Lattice Energy} - (\text{Hydration Energy of } K^+ + \text{Hydration Energy of } Cl^-)$$

In the given problem, we have:

- $\text{Hydration energy of } K^+ = -x\,\text{kJ/mol}$
- $\text{Hydration energy of } Cl^- = -y\,\text{kJ/mol}$
- $\text{Lattice energy of KCl} = -z\,\text{kJ/mol}$

Substitute these values into the formula:

$$\text{Heat of Solution} = -z - (-x + -y) = -z + (x + y)$$

Simplifying the expression, we get:

$$\text{Heat of Solution} = z - (x + y)$$

Thus, the correct option is:

$z-(x+y)$

Therefore, the heat of solution of KCl is calculated as the lattice energy minus the sum of the hydration energies of K^+ and Cl^-.

---

27. Answer: 2500 - 2500

Explanation:

Mass of n-octane = 1.14 g Heat capacity of the bomb calorimeter (C) = 5 kJ K^{-1} Increase in temperature (ΔT) = 5 K The heat evolved during the combustion of n-octane at constant volume (q_v) is absorbed by the calorimeter, causing the temperature increase.
Magnitude of heat evolved = $q_v = C \times \Delta T \ \ q_v = 5\,\text{kJ K}^{-1} \times 5\,\text{K} = 25\,\text{kJ}$
This is the heat evolved from the combustion of 1.14 g of n-octane. We need to find the heat of combustion per mole of n-octane.
The molecular formula of n-octane is C_8H_{18}.
The molar mass of n-octane is: $(8 \times 12) + (18 \times 1) = 96 + 18 = 114\,\text{g mol}^{-1}$ Number of moles of n-octane burnt = $\frac{\text{mass of n-octane}}{\text{molar mass of n-octane}}$

$$\text{Moles of n-octane} = \frac{1.14\,\text{g}}{114\,\text{g mol}^{-1}} = 0.01\,\text{mol}$$

The heat evolved from the combustion of 0.01 mol of n-octane is 25 kJ. The heat of combustion per mole of n-octane (ΔU) at constant volume is:

$$\Delta U = \frac{\text{Heat evolved}}{\text{Moles of n-octane}} = \frac{25\,\text{kJ}}{0.01\,\text{mol}} = 2500\,\text{kJ mol}^{-1}$$

The magnitude of the heat of combustion at constant volume is 2500 kJ mol^{-1}.
The nearest integer is 2500.

---

28. Answer: 304 - 304

Explanation:

The problem asks for the calculation of the total work done (w) when a perfect gas undergoes a series of reversible transformations as depicted in the given P-V diagram, from an initial point 1 to a final point 4.

Concept Used:

The work done (w) in a reversible thermodynamic process is calculated by the integral of pressure with respect to volume. The convention used in chemistry is:

$$w = -\int P_{ext}\, dV$$

For a reversible process, $P_{ext} = P_{gas} = P$. The total work done for a process consisting of multiple steps is the sum of the work done in each individual step: $w_{total} = w_{1\to2} + w_{2\to3} + w_{3\to4}$.

- For an **isochoric process** (constant volume), $dV = 0$, so the work done is $w = 0$.
- For an **isobaric process** (constant pressure), the work done is $w = -P\Delta V = -P(V_{final} - V_{initial})$.

The final answer needs to be converted from L·atm to Joules (J) using the conversion factor \(1 \text{ L·atm} = 101.325 \text{ J}\).

Step-by-Step Solution:

Step 1: Analyze and calculate the work done for the path from point 1 to point 2 ($w_{1\to2}$).

This is an isochoric process, as the volume remains constant at $V = 1000 \text{ cm}^3$. For any isochoric process, the change in volume dV is zero.

$$\Delta V = V_2 - V_1 = 1000 - 1000 = 0 \text{ cm}^3$$

$$w_{1\to2} = -P\Delta V = 0 \text{ J}$$

Step 2: Analyze and calculate the work done for the path from point 2 to point 3 ($w_{2\to3}$).

This is an isobaric process, as the pressure remains constant at $P = 3.00 \text{ atm}$. The volume changes from $V_2 = 1000 \text{ cm}^3$ to $V_3 = 2000 \text{ cm}^3$.

First, calculate the change in volume ΔV:

$$\Delta V = V_3 - V_2 = 2000 \text{ cm}^3 - 1000 \text{ cm}^3 = 1000 \text{ cm}^3$$

Convert the volume from cm^3 to Liters (L), since $1 \text{ L} = 1000 \text{ cm}^3$:

$$\Delta V = 1 \text{ L}$$

Now, calculate the work done in L·atm:

$$w_{2\to3} = -P\Delta V = -(3.00 \text{ atm})(1 \text{ L}) = -3.00 \text{ Latm}$$

Step 3: Analyze and calculate the work done for the path from point 3 to point 4 ($w_{3\to4}$).

This is also an isochoric process, with the volume constant at $V = 2000 \text{ cm}^3$.

$$\Delta V = V_4 - V_3 = 2000 - 2000 = 0 \text{ cm}^3$$

$$w_{3\to4} = -P\Delta V = 0 \text{ J}$$

Step 4: Calculate the total work done (w_{total}) for the entire transformation from 1 to 4.

The total work is the sum of the work done in each step.

$$w_{total} = w_{1\to2} + w_{2\to3} + w_{3\to4}$$

$$w_{total} = 0 + (-3.00 \text{ Latm}) + 0 = -3.00 \text{ Latm}$$

Step 5: Convert the total work from L·atm to Joules (J).

Using the conversion factor $1 \text{ Latm} = 101.325 \text{ J}$:

$$w_{total} = -3.00 \text{ Latm} \times 101.325 \frac{\text{J}}{\text{Latm}}$$

$$w_{total} = -303.975 \text{ J}$$

Step 6: Round the result to the nearest integer as requested by the question.

$$w_{total} \approx -304 \text{ J}$$

The question asks for the value to be filled in the blank of `(-)___ J`. Therefore, the value is the magnitude of the work done.

The total work done (w) is **304** J.

---

29. **Answer: c**

Explanation:

To solve this question, let's examine each statement individually and understand the underlying principles of each.

1. **Statement I:** When a system containing ice in equilibrium with water is heated, heat is absorbed by the system, and there is no change in the temperature of the system until the whole ice gets melted.
    - **Explanation:** This describes the process of melting, where the temperature remains constant as heat is absorbed. This heat is absorbed as latent heat of fusion, which is used to change the ice to water without increasing the temperature until all ice has melted.
    - **Conclusion:** Statement I is true.
2. **Statement II:** At the melting point of ice, there is an absorption of heat to overcome intermolecular forces of attraction within the molecules of water in ice, and the kinetic energy of molecules is not increased at the melting point.
    - **Explanation:** At the melting point, the absorbed heat breaks the solid lattice (overcoming intermolecular forces) rather than increasing the kinetic energy, which explains why temperature remains constant during this phase transition.
    - **Conclusion:** Statement II is true.

Thus, both statements correctly describe the thermodynamics involved in the melting of ice:

- **Correct Answer:** Both Statement I and Statement II are true

This choice is consistent with the principles of phase transition and heat absorption without temperature change (latent heat) during the melting process.

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30. Answer: b

Explanation:

The question involves understanding the thermodynamics of dissolution and dilution of hydrochloric acid (HCl) gas in water, represented by two reactions with their enthalpy changes (ΔH).

1. The process of dissolving a gas in water is generally exothermic, meaning it releases heat. That is why the enthalpy change (ΔH) is negative in both reactions given in the question.

2. The first option, "Dissolution of gas in water is an endothermic process," is incorrect because dissolution is exothermic (negative ΔH).
3. The second option, "The heat of solution depends on the amount of solvent," suggests that the amount of heat evolved or absorbed can vary with different amounts of solvent. This statement matches our data, as the heat of solution does change between the two cases (10 and $40\ H_2O$). Thus, this option is correct.
4. The third option involves calculating the heat of dilution from $HCl.10\ H_2O$ to $HCl.40\ H_2O$. The heat of dilution can be determined by calculating the ΔH for the change in the number of water molecules:

$$\Delta H_{\text{dilution}} = \Delta H(\text{HCl.40 H}_2\text{O}) - \Delta H(\text{HCl.10 H}_2\text{O})$$

$$\Delta H_{\text{dilution}} = -72.79\,\text{kJ/mol} - (-69.01\,\text{kJ/mol}) = -3.78\,\text{kJ/mol}$$

   Thus, this option is correct as well, since it calculates the heat of dilution correctly.
5. The fourth option claims both expressions represent heat of formation, but generally, they represent the heat of solution rather than formation. Therefore, it does not accurately describe the scenario provided.

Hence, the correct statement is: "The heat of solution depends on the amount of solvent," highlighting the impact of the volume of solvent on the enthalpy change during solution formation.