Thermochemistry Chemistry Questions with Solutions

Q-1: Given that standard molar enthalpies of formation of NO(g) and NO_2(g) are respectively 90.3 kJ/mol and 33.2 kJ/mol. Calculate the enthalpy change for the reaction $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$.

Answer: $\Delta H = \sum \Delta H_{products} - \sum \Delta H_{reactants}$
$\Delta H = 2(33.2)-[2(90.3)+0]$
$\Delta H = -114.2kJ$

Note: Enthalpy of formation of free elements like O_2 is always 0.

Q-2: The enthalpy of formation of AgCl is obtained from the enthalpy change from which one of the following process?

a) $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$
b) $Ag(s) + ½ Cl_2(g) \rightarrow AgCl(s)$
c) $AgCl(s) \rightarrow Ag(s) + ½ Cl_2(g)$
d) $Ag(s) + AuCl \rightarrow Au(s) + AgCl(s)$

Answer: b) $Ag(s) + ½ Cl_2(g) \rightarrow AgCl(s)$

Explanation: The enthalpy change accompanying the formation of one mole of a pure substance in its standard state from the constituents elements in their standard state is called enthalpy of formation. For Ag, the standard state is solid and for chlorine, it is gas. Thus, the process that gives the enthalpy of formation of AgCl is: $Ag(s) + ½ Cl_2(g) \rightarrow AgCl(s)$

Q-3: ΔH and ΔE for the reaction, $Fe_2O_3 (s) + 3H_2(g) \rightarrow 2Fe(s) + H_2O(l)$ at constant temperature are related as

a) $\Delta H = \Delta E$
b) $\Delta H = \Delta E + RT$
c) $\Delta H = \Delta E + 3RT$
d) $\Delta H = \Delta E - 3RT$

Answer: d) $\Delta H = \Delta E - 3RT$

Explanation: For any chemical reaction,

$\Delta H = \Delta E + \Delta_{ng}RT$
Where, Δ_{ng} = total number of moles of gaseous product- total number of moles of gaseous reactants.

For the given reaction,
Δ_{ng} =0-3 = -3

$\Delta H = \Delta E + (-3)RT$
$\Delta H = \Delta E - 3RT$

Q-4: For the reaction, $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$, compute the entropy change (in J/K/mol) for the process and comment on the sign of the property.

Data	Species	$NH_3(g)$	$N_2(g)$	$H_2(g)$
	S°(J/K/mol)	192.3	191.5	130.6

Answer: $\Delta S^o = \sum \Delta S^o_{products} - \sum \Delta S^o_{reactants}$

$\Delta S^o = 2\times192.3-[191.5+3(130.6)] = -198.7 J/K/mol$

In the reaction, two gaseous species are there in the reactant side and the number of moles of gaseous species on the product side is 1. So there is a decrease in the number of gaseous species which is represented by the negative sign of entropy.

Q-5: The rate of evaporation of a liquid is always faster at a higher temperature because

a) The enthalpy of vaporisation is always endothermic
b) The enthalpy of vaporisation is always exothermic
c) The enthalpy of vaporisation is zero
d) The internal pressure of the liquid is less than that of the gas

Answer: a) The enthalpy of vaporisation is always endothermic

Explanation: At the time of evaporation, energy is required to overcome the intermolecular forces between the molecules of liquid, that is, the enthalpy of vaporisation is always endothermic. Since at high temperatures more energy is available, the rate of evaporation is faster at elevated temperatures.

Q-6: Given that

$A(s) \rightarrow A(l) \quad \Delta H = +x$
$A(l) \rightarrow A(g) \quad \Delta H = +y$

The heat of sublimation of A will be

a) x-y
b) x+y
c) x or y
d) -x+y

Answer: b) x+y

Explanation: It is a physical process by which a substance is directly converted from solid to gaseous.

The reaction for heat of sublimation is A(s) → A(g) which is obtained by adding A(s) → A(l) and A(l) → A(g). Thus, the heat of sublimation of A will be x+y.

Q-7: ΔH= -25kcal for the reaction $CH_4(g) + Cl_2(g) \rightarrow CH_3Cl(g) + HCl(g)$.

Bond	Bond Energy(kCal)
C-Cl	84
H-Cl	103
C-H	x
Cl-Cl	y
x:y	9:5

What is the bond enthalpy of Cl-Cl bond? Use the given data to calculate the answer.

Answer: $\Delta H = \sum \Delta B.E_{reactants} - \sum \Delta B.E_{products}$
ΔH= 4(C-H)+Cl-Cl-[3(C-H)+1(C-Cl) +1(H-Cl)]
-25= 4x+y-[3x+84+103]

This implies, x+y-187= -25
Or x+y= 162

Given: x:y=9:5. Therefore, x= 9y/5

Substituting x= 9y/5 in x+y= 162 , we get, y= 57.85 kCal

Q-8: For the reaction,
$N_2(g) + 2O_2(g) \rightarrow 2NO_2(g)$ ΔH = -66 kJ

Calculate the value of Δ_fH of NO_2.

Answer: Enthalpy of formation(Δ_fH) is the enthalpy change accompanying the formation of one mole of a pure substance in its standard state from the constituent elements in their standard state.

The given ΔH of NO_2 is for its two mole formation. Since we need to calculate the Δ_fH of NO_2 for one mole by definition, therefore it will be equal to
$\Delta_fH = \Delta H/2 = -66/2 = -33$kJ

Q-9: Which of the following is not an endothermic reaction?

a) Decomposition of water
b) Conversion of graphite to diamond
c) Combustion of methane
d) Dehydrogenation of ethane to ethylene

Answer: c) Combustion of methane

Explanation: Heat is produced in any combustion reaction. The reactions that involve the release of heat are said to be exothermic.

Q-10: The enthalpy changes of the following reactions at 27°C are

1) $Na(s) + \frac{1}{2} Cl_2(g) \rightarrow NaCl(s)$; $\Delta_fH = -411$kJ/mol
2) $H_2(g) + S(s) + 2O_2(g) \rightarrow H_2SO_4(l)$; $\Delta_fH = -811$kJ/mol
3) $2Na(s) + S(s) + 2O_2(g) \rightarrow Na_2SO_4(s)$; $\Delta_fH = -1382$kJ/mol
4) $\frac{1}{2} H_2(g) + \frac{1}{2} Cl_2(g) \rightarrow HCl(g)$; $\Delta_fH = -92$ kJ/mol
From the above data, the heat change of reaction at constant volume (in kJ/mol) at 27°C for the process
$2NaCl(s) + H_2SO_4(l) \rightarrow Na_2SO_4(s) + 2HCl(g)$ is:

Answer:
Reverse equation 1 and multiply it by 2
a) $2NaCl(s) \rightarrow 2Na(s) + Cl_2(g)$; $\Delta_fH = +2\times411$kJ/mol = +822kJ/mol
Reverse equation 2
b) $H_2SO_4(l) \rightarrow H_2(g) + S(s) + 2O_2(g)$; $\Delta_fH = +811$kJ/mol
c) $2Na(s) + S(s) + 2O_2(g) \rightarrow Na_2SO_4(s)$; $\Delta_fH = -1382$kJ/mol
Multiply equation 4 by 2
d) $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$; $\Delta_fH = -92\times2 = -184$ kJ/mol

On adding equation, a, b,c,and d, we get

$2NaCl(s) + H_2SO_4 (l) \rightarrow Na_2SO_4 (s) + 2HCl (g); \Delta_fH$

$\Delta_fH = 822+811+(-1382-184) = 67$kJ/mol

We know that, $\Delta_fH = \Delta U + \Delta n_gRT$

Where, Δ_{ng} = total number of moles of gaseous product- total number of moles of gaseous reactants.
For the reaction given, $\Delta_{ng} = 2$
R= 8.3J/K mol = 0.0083kJ/mol
T= 27°C = 300K
ΔU is the heat change of reaction at constant volume
On substituting the values, we get
$67 = \Delta U + 2\times0.0083\times300$

ΔU= 62.02 kJ/mol

Q-11: Heat of combustion for benzene and acetylene are -3900 and -642 joule. Then calculate the heat of reaction(per mole) for the following reaction.
$3C_2H_2 \rightarrow C_6H_6$

Answer:
1) $C_6H_6 + 15/2\ O_2 \rightarrow 6CO_2 + 3H_2O$
2) $C_2H_2 + 5/2\ O_2 \rightarrow 2CO_2 + H_2O$
Multiply reaction 2 with 3 and reverse reaction 1 and add them, we get
$3C_2H_2 \rightarrow C_6H_6$

$\Delta_fH = 3(\Delta H_C)_{acetylene} + (\Delta H_c)_{benzene} = 3(-642) + 3900 = 1974$ J for 3 mole
Therefore, Δ_fH for 1 mole= 1974/3 = 658J/mol

Q-12: For combustion of one mole of magnesium in an open container at 300 K and 1 bar pressure, $\Delta_CH = -601.70$kJ/mol, the magnitude of change in internal energy for the reaction is _________ kJ. (Nearest integer)
Given: 8.3 J/Kmol

Answer: The combustion reaction for one mole of magnesium is
$Mg(s) + ½\ O_2(g) \rightarrow MgO(s)$

$\Delta_CH = \Delta E + \Delta_{ng}RT$
Where, Δ_{ng} = total number of moles of gaseous product- total number of moles of gaseous reactants.
ΔE is the internal energy
For the given reaction, $\Delta_{ng} = -½$

ΔE= -601.70 kJ/mol -½ (8.3×10^{-3} × 300) = 599.455 kJ = 600 kJ

Q-13: NaOH(s) has a heat of solution of -42.6kcal/mol NaOH. When NaOH is dissolved in water, the temperature of the solution

a) increases
b) decreases
c) remains constant
d) can't be predicted

Answer: a) increases

Q-14: What is Hess's Law?

Answer: Hess's law is an important result of the first law of thermodynamics. It states that the enthalpy change in a chemical or physical process is the same whether it is performed in one step or several steps.

Q-15: If total enthalpies of reactants and products are H_R and H_P respectively, then for an endothermic reaction

a) $H_R<H_P$
b) $H_R>H_P$
c) $H_R=H_P$
d) Data insufficient

Answer: a) $H_R<H_P$

Explanation: For an endothermic reaction, enthalpy of reaction is positive. Enthalpy of reaction is calculated by using below equation:

$\Delta H= \sum\Delta H_P- \sum\Delta H_R$

For ΔH to be positive, H_R must be less than H_P.

Practise Questions on Thermochemistry

Q-1: Which of the following reactions does not represent $\Delta_f H$?

a) $C(s) + O_2 \rightarrow CO_2$
b) $Br_2(l) + H_2(g)\rightarrow 2HBr$
c) $CO + O_2 \rightarrow CO_2$
d) $C(s)\rightarrow C(g)$

Answer: b) and c)

Q-2: Which of the following pairs are correctly matched?

i) Arrhenius Equation	Variation of enthalpy of a reaction with temperature
ii) Kirchhoff equation	Variation of rate constant with temperature
iii) Second law of thermodynamics	Entropy of an isolated system tends to increase and reach a maximum value
iv) Hess's law of constant heat summation	Enthalpy change in a reaction is always constant and independent of the manner in which the reaction occurs.

Answer: iii and iv

Explanation: Arrhenius equation is the variation of rate constant with temperature. On the other hand, the Kirchhoff equation explained the temperature dependence of enthalpy of reactions.

Q-3: From the following data at 25°C

Reaction	$\Delta_f H^\circ$ (kJ/mol)
$\frac{1}{2} H_2(g) + \frac{1}{2} O_2(g) \rightarrow OH(g)$	42
$H_2(g) + \frac{1}{2} O_2(g) \rightarrow H_2O(g)$	-242
$H_2(g) \rightarrow 2H(g)$	436
$O_2(g) \rightarrow 2O(g)$	495

Calculate the $\Delta_f H^\circ$ for the reaction $H_2O(g) \rightarrow 2H(g) + O(g)$

Answer: On reversing the equation $H_2(g) + \frac{1}{2} O_2(g) \rightarrow H_2O(g)$, we get

$H_2O(g) \rightarrow H_2(g) + \frac{1}{2} O_2(g)$; $\Delta_f H^\circ = + 242$ kJ/mol …………1)

Multiply ½ to the equation $O_2(g) \rightarrow 2O(g)$, we get

$\frac{1}{2} O_2(g) \rightarrow O(g)$; $\Delta_f H^\circ = 495/2$ kJ/mol ………….2)
$H_2(g) \rightarrow 2H(g)$; $\Delta_f H^\circ = 436$ kJ/mol………….3)

On adding, 1), 2) and 3), we get

$H_2O(g) \rightarrow O(g) + 2H(g)$ …$\Delta_f H^o$

$\Delta_f H^o = (242 + 495/2 + 436)$ kJ/mol $= 925.5$ kJ/mol

Q-4: The heat of combustion of carbon is 394 kJ/mol. The heat evolved in combustion of 6.022×10^{22} atoms of carbon is:

Answer: The reaction for the combustion of carbon is
$C(s) + O_2(g) \rightarrow CO_2(g)$ $\Delta H = 394$ kJ/mol

From the equation, we can say that
Heat of combustion for 1 mol of carbon = 394 kJ/mol
1 mol of carbon contains 6.022×10^{23} atoms of carbon

It means,
Heat of combustion for 6.022×10^{23} atoms of carbon = 394 kJ/mol

Heat of combustion for 6.022×10^{22} atoms of carbon = $(394 \times 6.022 \times 10^{22})/6.022 \times 10^{23}$
= 39.4 kJ/mol

Q-5: Calculate standard enthalpies of formation of $CS_2(l)$. Given the standard enthalpy of combustion of carbon (s), sulphur(s) and $CS_2(l)$ are: -393.3, -293.72 and -1108.76 kJ/mol respectively.

Answer:
The combustion reaction for Carbon is $C(s) + O_2(g) \rightarrow CO_2(g)$ $\Delta H = -393.3$ kJ/mol
The combustion reaction for sulphur is $S(s) + O_2(g) \rightarrow SO_2(g)$ $\Delta H = -293.72$ kJ/mol
The combustion reaction for $CS_2(l)$ is $CS_2(l) + 3O_2(g) \rightarrow 2SO_2(g) + CO_2(g)$ $\Delta H = -1108.76$ kJ/mol

Reverse the reaction for the combustion reaction for $CS_2(l)$, we get
$2SO_2(g) + CO_2(g) \rightarrow CS_2(l) + 3O_2(g)$ $\Delta H = +1108.76$ kJ/mol …….1)

Multiply the combustion reaction for sulphur by 2, we get
$2S(s) + 2O_2(g) \rightarrow 2SO_2(g)$ $\Delta H = -2 \times 293.72 = -587.44$ kJ/mol …….2)

$C(s) + O_2(g) \rightarrow CO_2(g)$ $\Delta H = -393.3$ kJ/mol ……..3)

On adding, 1+2+3, we get,
$C(s) + 2S(s) \rightarrow CS_2(l)$

Above equation represents the formation reaction of $CS_2(l)$ from its constituent elements in their standard states. Thus, $\Delta_f H(CS_2(l)) = +1108.76 - 587.44 - 393.3 = 128.02$ kJ/mol