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Class 11
Chemistry

Chapter 6
Thermochemistry
Handwritten Notes

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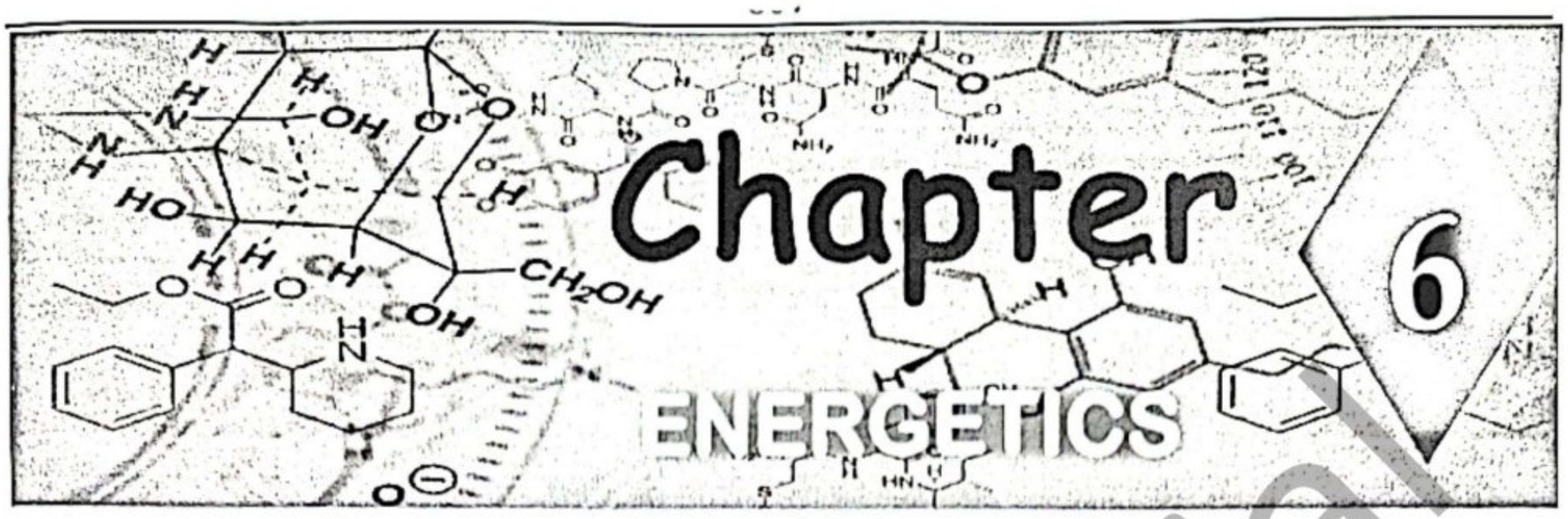


INTRODUCTION:

Energetics

The study of all types of energy changes associated with physical and chemical changes is known as **energetics**

Thermo chemistry:

The branch of chemistry which deals with the study of heat changes during the chemical reaction is called thermochemistry.

It deals with the:

- Principles and laws of energy changes
- The role of energy in chemical reactions
- Measurement of heat of chemical reaction.

ENERGY IN CHEMICAL REACTIONS:

i. The energy in the form of heat is either evolved or absorbed as a result of chemical reaction. This is because, in chemical reaction old bonds are broken and new bonds are formed.

ii. Bond breaking always consumes energy and bond making always releases energy.

iii. When the energy released by bond forming is greater than the energy consumed by bond breaking, there is a net release of chemical energy.

iv. Whereas energy is absorbed, when the energy of bond breaking is greater than the energy of bond forming. Thus, in chemical reactions energy is exchanged with the surroundings.

Units of Thermal energy

Unit of heat or thermal energy used in SI system is the Joule (J).

It is defined as the energy required in moving an object or a particle through a distance of one meter by a force of one Newton.

Another common unit of heat is **calorie.** 1 Calorie = 4.184 Joules

It is defined as heat required to raise the temperature of one gram of water by 1 C^0

Definitions of Terms used in Thermodynamics

1. **System:**

The part of the universe which is under observation is called a system.

2. **Surroundings:**

The part of the universe except system is called surroundings.

Boundary:

The real or imaginary surface separating the system from surroundings is called boundary.

It may be real or imaginary.

e.g. ice melts in a beaker.

In this case, ice is a system and beaker and other things are surrounding.

THERMOCHEMICAL REACTIONS:

Reactions in which heat is released or absorbed, are called thermochemical reactions, and equations obtained are called thermochemical Equations.

TYPES OF CHEMICAL REACTIONS

(i) **Exothermic Reactions:**

(ii) **Endothermic Reactions:**

Exothermic Reactions	Endothermic Reaction
1. The chemical reaction that proceeds with evolution of heat is called exothermic reaction.	1. A chemical reaction that proceeds with the absorption of heat is called an endothermic reaction.
2. In this reaction, heat transfers from system to surrounding.	2. In this reaction heat is transferred from surrounding to the system.
3. The container becomes hot in which reaction is taking place.	3. The container becomes cool in which reaction is taking place.
4. In these reactions enthalpy of products is less than that of reactants.	4. In these reactions enthalpy of products is more than that of reactants.
5. Dissolving of calcium oxide in water is an example of exothermic reaction. $CaO + H_2O \longrightarrow Ca(OH)_2 + Heat$ $C(s) + O_2(g) \longrightarrow CO_2(g)\ \Delta H_r^o = -393.5KJ$ $2H_2(g) + O_{2(g)} \longrightarrow 2H_2O(l)\ \Delta H_r^o = -571.6KJ$	5. For examples $H_2(g) + I_2(g) \longrightarrow 2HI(g)\quad \Delta H_r^o = +53.8KJ$ $C(s) + H_2O(g) \longrightarrow CO(g) + H_2(g)\ \Delta H_r^o = +131.4KJ$ $N_2(g) + O_2(g) \longrightarrow 2NO(g)\quad \Delta H_r^o = +180.5KJ$

Concept Assessment Exercise 6.1

Classify the following processes as exothermic or endo thermic.

a) Freezing of water

Freezing occurs with the loss of heat to the surrounding; therefore, it is an exothermic process.

b)Combustion of methane

During combustion process, heat is released to the surrounding, therefore, it is a exothermic process.

c)Sublimation of dry ice:

The conversion of solid into vapours is called sublimation. To convert solid into vapours energy is require, therefore, it is an endothermic process.

d)$H_2O_{(g)} \rightarrow H_2O_{(l)}$

When gaseous water is cooled, it is converted into liquid water. Thus, energy is released to the surrounding, therefore, it is an exothermic process.

e)Decomposition of limestone:

To decompose a substance, energy is required, therefore, it is an endothermic process.

Heat of Reaction (ΔH_r^o)

Amount of heat, released or absorbed, when no. of moles of reactant react with each other and form product, under standard conditions is called standard enthalpy of reaction. All the reactions and product must be in their standard physical state.

e.g. $2H_{2(g)} + O_{2\,(g)} \longrightarrow 2H_2O\,(l)$ $\quad \Delta H_r^o = -571.6KJ$

Relationship between Enthalpy Change and Heat of Reaction or Heat of Combustion of a Reaction.

Because most chemical reactions occur at constant pressure, we can equate the heat change in these reactions to the change in enthalpy. This means we can define heat of a reaction as the change in enthalpy (ΔH), as the difference between the enthalpies of the products (final estate) and the enthalpies of the reactants (initial state).

$$\Delta H = H_{products} - H_{reactants}$$

The enthalpy changes of a reaction can be positive or negative depending upon the process, for an exothermic process, in which heat is released by the system of the surrounding AH is negative for an endothermic process, in which heat is absorbed by the system from the surroundings ΔH is positive Since, combustion is an exothermic process in which heat is released by the system to its surroundings, ΔH is negative.

Standard states and standard enthalpy changes

As ΔH varies with conditions, we use standardized ΔH values. These values are calculated when all the substances are in their standard state.

Conditions for the standard states are as follows:

1.Standard state for a gas is 1 atm.

2.Standard state for an element or a compound is the most stable physical state at 1atm and 25°C (298 K).

3.Standard state for a substance in aqueous solution is 1M concentration.

Define the following terms.

1. Standard Enthalpy of Reaction (ΔH_r^o)

Amount of heat, released or absorbed, when no. of moles of reactant react with each other and form product, under standard conditions is called standard enthalpy of reaction.

All the reactions and product must be in their standard physical state.

e.g. $2H_{2(g)} + O_{2\,(g)} \longrightarrow 2H_2O\,(l)$ $\quad \Delta H_r^o = -571.6KJ$

2. Standard Enthalpy of Formation: (ΔH_f^o)

Amount of heat, released or absorbed, when no. of moles of reactant react with each other and form one mole of product, under standard conditions, is called standard enthalpy of formation.

All the reactions and product must be in their standard physical state.

e.g. $Mg_{(s)} + \frac{1}{2}O_{2(g)} \longrightarrow MgO_{(s)}$ $\quad \Delta H_f^0 = -601.4kJ/mol$

Example 6.1

Write thermochemical equation for the formation of H_2O, SO_3 and H_2O_2.

Solution:

(i) $H_{2(g)} + ½\ O_{2(g)} \longrightarrow H_2O_{(l)}$ $\quad \Delta Hf = -285.8\ kJ\ mole^{-1}$

(ii) $S_{(s)} + \frac{3}{2}O_{2(g)} \longrightarrow SO_{3(g)}$ $\quad \Delta Hf = -395.2\ kJ\ mole^{-1}$

(iii) $H_{2(g)} + O_{2(g)} \longrightarrow H_2O_{2(l)}$ $\quad \Delta Hf = -191.2\ kJ\ mole^{-1}$

3. Standard Enthalpy of Combustion (ΔH_c^o)

Amount of heat released, when one mole compound is burn in free supply of air (O_2), under standard conditions, is called standard enthalpy of combustion.

e.g. $C_2H_5OH_{(\ell)} + 3O_{2(g)} \longrightarrow 2CO_{2(g)} + 3H_2O_{(g)} \quad \Delta H_c^0 = -1365 kJ/mol$

Example 6.2

Write thermochemical equation for the combustion of, C, CH_4 and H_2.

Solution:

(i) $C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)}$ $\quad \Delta Hc^o = -393.5 \text{ kJ mole}^{-1}$

(ii) $CH_{4(g)} + 2O2(g) \longrightarrow CO2(g) + 2HO(\ell)$ $\quad \Delta Hc^o = -890.4 \text{ kJ mole}^{-1}$

(iii) $H_{2(g)} + \frac{1}{2}O_{2(g)} \longrightarrow H_2O_{(\ell)}$ $\quad \Delta Hc^o = -285.8 \text{ kJ mole}^{-1}$

4. Standard Enthalpy of Atomization (ΔH_{at}^o)

The enthalpy change when one mole of gaseous atoms are formed from its element under standard

e.g. (i) $\frac{1}{2} H_2 (g) \longrightarrow H (g)$ $\quad \Delta H_{at}^o = +218 \text{ KJ mole}^{-1}$

(ii) $\frac{1}{2} Cl_2 (g) \longrightarrow Cl (g)$ $\quad (\Delta H_{at}^{\bullet}) = +121 \text{ KJ mole}^{-1}$

5. Standard Enthalpy of Neutralization (ΔH_n^o)

It is defined as the amount of heat evolved when one mole of H^+ ions from an acid combines with one mole of OH^- ions from a base to form one mole of water under standard conditions.

e.g.

$NaOH_{(aq)} + HCl_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)} \quad (\Delta H_n^{\bullet}) = -57.4 \text{KJ mole}^{-1}$

It is same for strong acid and base because net chemical change is formation of one mole of water

6. Standard Enthalpy of Solution (ΔH_{sol}^o)

It is the enthalpy change when one mole of a substance is dissolved in so much solvent that further dilution results in no detectable heat change, under standard conditions.

e.g. $NH_4Cl\ (s) \xrightarrow{H_2O} NH_4^{+1}\ (aq) + Cl\text{-}(aq) \quad (\Delta H_{sol}^{\bullet}) = +15.1 \text{KJ mole}^{-1}$

7. Standard Enthalpy of first Electron Affinity ($\Delta H_{E.A}$)

Amount of heat released when one mole of gaseous atom gain an electron to form one mole of gaseous ion Under standard conditions

$Cl_{(g)} + e^- \longrightarrow Cl^-_{(g)} \quad \Delta H_{E.A}^o = -349 \text{ KJ/mol}$

8. Lattice energy (ΔH_{latt}^o)

Amount of energy released when one mole ionic compound is formed from its opposite charged ions in gaseous state.

e.g $Na^+_{(g)} + Cl^-_{(g)} \longrightarrow -NaCl_{(s)} \quad \Delta H_r = -787 kJ/mol$

Bond Energy

Bond Dissociation Energy:

The amount of energy required to break one mole of a particular bond to form neutral atoms is called bond dissociation energy.

Bond Energy:

The amount of energy released when one mole of a particular bond forms from neutral atoms is called bond energy.

When a chemical reaction occurs, old bonds are broken and new bonds are formed. Breaking bonds always requires energy, and forming a bond always releases energy. The difference between bond dissociation energy and bond energy determines whether the reaction absorbs or releases energy overall. For any chemical reaction, the enthalpy change is the sum of the bond dissociation energies of the reactants minus the sum of the bond energies of the products.

Example 6.3

We can theoretically calculate bond energy of a bond from the known bond energy data. Example, we can calculate bond energy on H—Cl as follows;

Bond dissociation energy of H_2 = 436 kJ/mol.

Bond dissociation energy of Cl_2 = 243 kJ/mol.

Bond energy of HCl = ?

Total energy absorbed in bond breaking = Total energy released in bond formatio

Bond dissociation energy of H—H + Bond dissociation energy of Cl — Cl = 2 (Bond energy of H-Cl)

436 + 242 =2 (Bond energy of H- Cl)

679 = 2 (Bond energy of H -Cl)

Bond energy of H — Cl = 339 kJ/mol.

Thus theoretically calculated bond energy of HCl = 339 kJ/mol.

Experimentally determined bond energy of HCl = 431.5 kJ/mol.

This is theoritcal values are average. These donot take into many factors like electronegativity differences of atoms.For any chemical reaction, the enthalpy change is the sum of bond dissociation energies of the reactants minus the sum of bond energies of products.

$\Delta H = \Sigma H$ (bonds broken) $- \Sigma H$ (bonds formed)

For the reaction between H_2 and O_2 to produce H_2O.

$2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$ ΔH = ?

ΔH^0 = 2(B.D.E of $H_2(g)$) + B.D.E of $O_2(g)$ — 4(B.E of O -H bonds)

= 2(436 kJ) + 493.6 kJ — 4(460 kJ)

= 872 kJ + 493.6 kJ — 1840 kJ

= –474.4 kJ

Thus the reaction between hydrogen and oxygen to form water is exothermic.

Variability in Bond Energies

Some bond energies are precise, others are approximate.

1. Diatomic molecules (H_2, O_2, N_2, Cl_2) have relatively accurate bond energies.
2. Complex molecules have more approximate bond energies due to factors like neighboring atoms and molecular geometry.
3. Bond strengths vary in different compounds (e.g., O-H in methanol vs. water).

HEAT CAPACITY

Define (a) Heat capacity (b) Specific heat capacity (c) Molar heat capacity

It is define as;

"The amount of heat required to raise the temperature of given amount of a substance by 1 Kelvin is called heat capacity."

The different forms of energy can be converted into one another. Any other form of energy can be completely transformed into heat energy. But at constant temperature heat cannot be completely converted into another form of energy. Temperature is not the same as heat.

Temperature difference determines the direction in which heat flows spontaneously. But quantity of heat within a substance at a given temperature is directly proportional to its mass. Therefore, amount of heat absorbed by a substance is proportional to the temperature change.

$$q \propto \Delta T$$

$$q = C \times \Delta T$$

Specific heat capacity:

"The amount of heat required to raise the temperature of given amount of a substance by 1Kelvin is called heat capacity."

$$\text{Specific heat capacity} = \frac{\text{Heat}}{\text{gram} \times \Delta T}$$

Molar heat capacity:

The amount of heat required to raise the temperature of one mole of a substance by one Kelvin is called its molar heat capacity.

$$\text{Molar heat Capacity} = \frac{\text{Heat}}{\text{moles} \times \Delta T}$$

Calorimetry:

It is the science of measuring heat of a chemical reaction by measuring the temperature change is called Calorimetry. Calorimeters measure the heat released from a system either at constant pressure ($q_p = \Delta H$) or at constant volume ($q_v = \Delta E$).

Thus there are two types of calorimetry.

i. Constant pressure Calorimetry
ii. Constant Volume Calorimetry

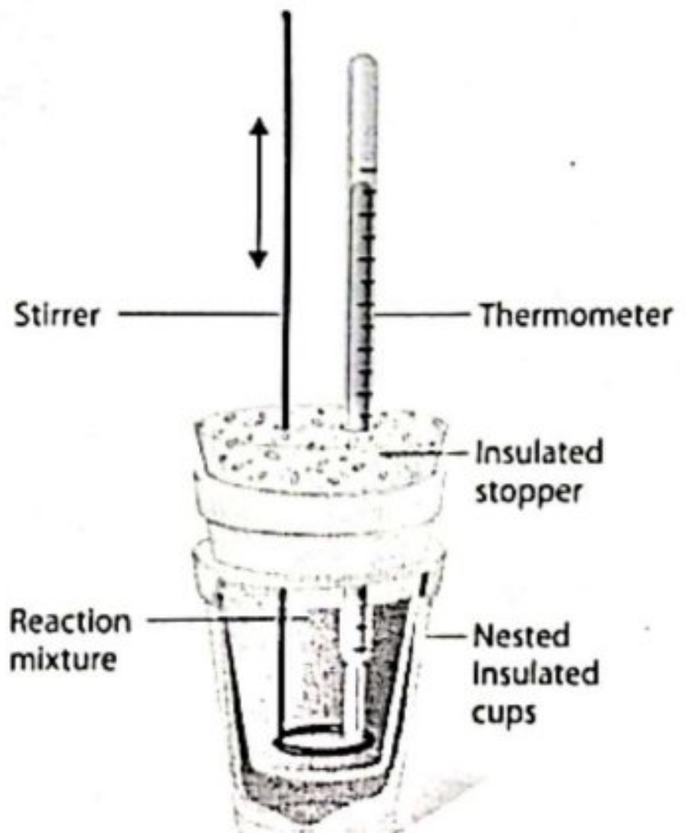
1. **Constant Pressure Calorimetry:**

In constant pressure calorimetry pressure of the system is fixed. For this purpose we need a thermally insulated container with a thermometer and stirrer. For most purposes a coffee cup calorimeter is used (Fig).

Coffee cup calorimeter

$$q = m \times c \times \Delta T$$

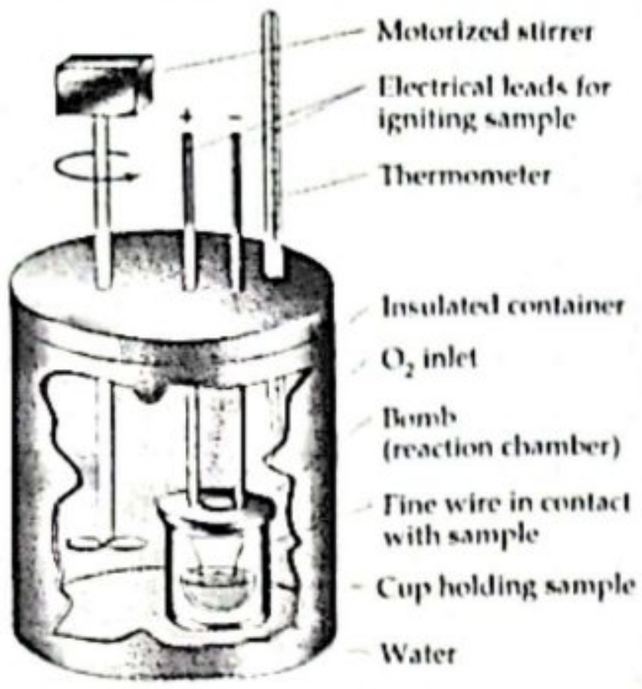
Where
m =mass of reactants
c=specific heat of reaction mixture
ΔT= change in temperature.



2. **Constant Volume Calorimetry**

The technique in which heat energy is calculated in a calorimeter at constant volume is Constant volume calorimetry. This technique is used for accurate determination of the enthalpy of

combustion for food, fuel and other compounds. A bomb calorimeter is used for this purpose. Chemical reaction in a bomb calorimeter takes place under constant volume conditions. A bomb calorimeter is shown in figure. It consists of a strong closed vessel (the bomb) immersed in an insulated water bath.



Hess's Law of constant Heat Summation:

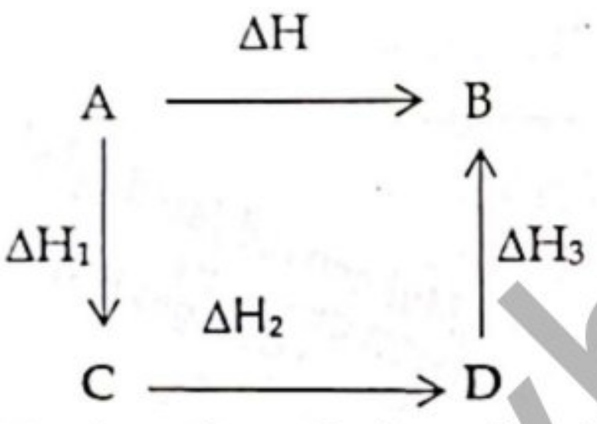
Q. Demonstrate Hess's law with the help of two suitable examples.

Ans. G.H. Hess stated this law in 1840.

According to this law, Enthalpy change during the chemical reaction, whether the reaction takes place in one step or more than one step.

Mathematically:

$$\Delta H = \Delta H_1 + \Delta H_2 + \Delta H_3$$



Explanation: Carbon dioxide can be prepared by two methods;

One-Step Method:

$$C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)} \quad \Delta H = -393.5KJ$$

Two-Step Method:

$$C_{(s)} + \frac{1}{2} O_{2(g)} \longrightarrow CO_{(g)} \qquad \Delta H_1 = -110.52KJ$$

$$CO_{(g)} + \frac{1}{2} O_{2(g)} \longrightarrow CO_{2(g)} \qquad \Delta H_2 = -282.98KJ$$

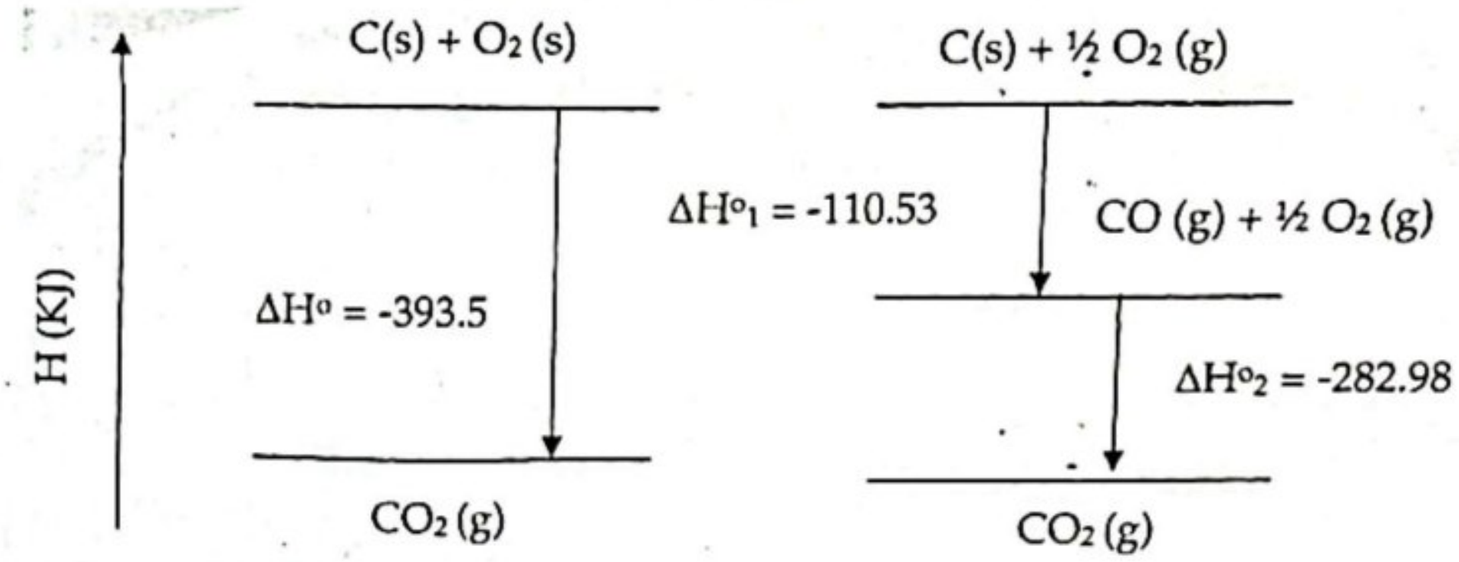
According to Hess's Law;

$\Delta H = \Delta H_1 + \Delta H_2$

$\Delta H = -110.52 - 282.98$

$\Delta H = -393.51$ kJ/mol

Energy cycle are



Importance of Hess's Law:

- These are many compounds which cannot be prepared directly from their elements. Some of these compounds cannot be decomposed into their constituent elements e.g. CCl_4. So, its enthalpy of formation cannot be determined directly by calorimetry.
- Some elements do not burn completely due to the formation of a protective covering on their surface. Such as Al, B etc. Thus, enthalpies of formation of Al_2O_3, B_2O_3 cannot be determined directly by calorimetry.
- Hess's Law is particularly useful for determining enthalpies of formation of such compounds.

Example 6.6

Combustion of C to CO2 evolves 393.5 KJ of energy.

$C_{(s)} + O_{2(g)} \longrightarrow 2CO_{2(g)}$ $\quad \Delta H = -393.5kJ$

This reaction may take place in two steps.

$C_{(s)} + O_{2(g)} \longrightarrow CO_{(g)}$ $\quad \Delta H_1 = -110.52kJ$

$CO_{(g)} + O_{2(g)} \longrightarrow CO_{2(g)}$ $\quad \Delta H_2 = -282.98kJ$

Net reaction $C_{(s)} + O_{2(g)} \longrightarrow CO_{(g)}$ $\quad \Delta H_1 + \Delta H_2 = -393.5kJ$

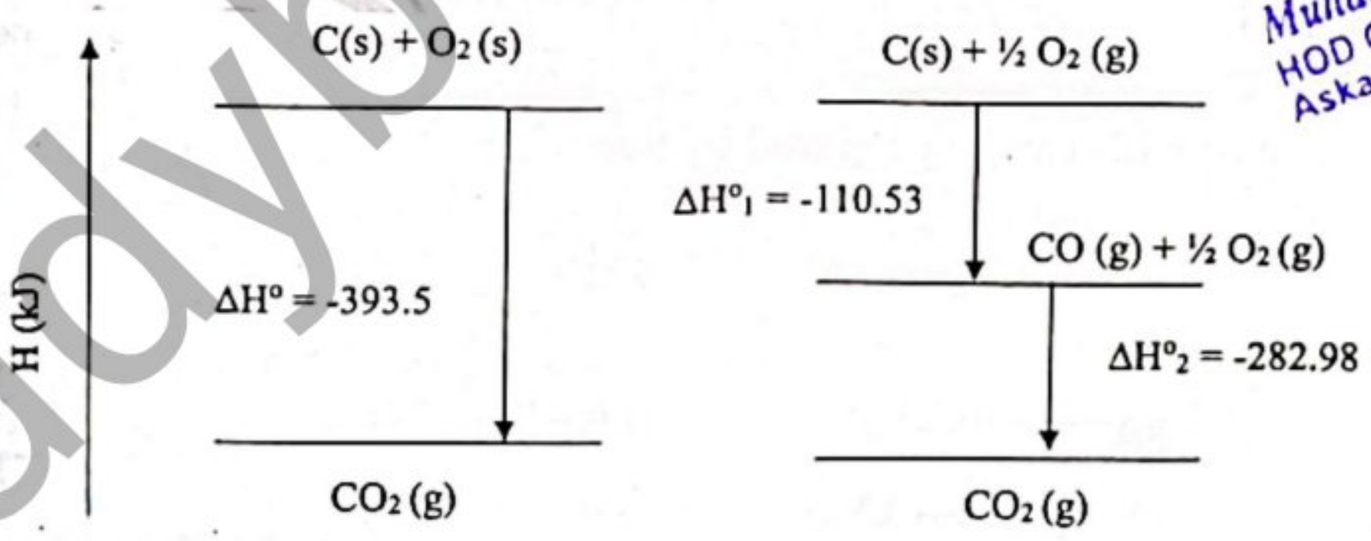


Figure 6.4 Energy cycle for the reaction between $C_{(s)}$ and $O_{2(g)}$ to produce CO_2 (g) which steps in this cycle is/are exothermic?

Example 6.7

Enthalpy of formation of methane cannot be measured directly. By the application of Mess's law it can be determined indirectly from the enthalpies of combustion for CH_4, H_2 and C.

(i) $CH_{4(g)} + 2O_{2(g)} \longrightarrow CO_{2(g)} + 2H_2O_{(\ell)}$ $\quad \Delta H° = -890.4\ kJ\ mole^{-1}$

(ii) $H_{2(g)} + \frac{1}{2}O_{2(g)} \longrightarrow H_2O_{(\ell)}$ $\Delta H° = -285.5\ kJ\ mole^{-1}$

(iii) $C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)}$ $\Delta H° = -393.5 kJ mole^{-1}$

(iv) $C_{(s)} + 2H_{2(g)} \longrightarrow CH_{4(g)}$ $\Delta H° = ?$

Equation (iii) has $C_{(S)}$ as reactant which is needed in the required equation. Thus equation (iii) will be used as such. Equation (ii) has $H_{2(g)}$ as reactant but the required equation needs $2H_{2(g)}$, thus equation (ii) will be used after multiplying by 2. Equation (i) has $CH_4(_g)$ as reactant, but this is needed as product in the required equation. Thus equation (i) must be reversed and the sign of $\Delta H°$ changed according. Adding the equations and deleting the species that occur on both sides we get.

$CO_{2(g)} + 2H_{2(g)}\ O_{(\ell)} \longrightarrow CH_{4(g)} + 2O_{(g)}$ $\Delta H° = +890.4\ kJ\ mole^{-1}$

$2H_{2(g)} + O2_{(g)} \longrightarrow 2H_2O_{(\ell)}$ $\Delta H° = -571.6\ kJ$

$2H_{2(g)} + O2_{(g)} \longrightarrow 2H_2O_{(\ell)}$ $\Delta H° = -571.6\ kJ$

$C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)}$ $\Delta H° = -393.5\ kJ$

$C_{(s)} + 2H_{2(g)} \longrightarrow CH_{4(g)}$ $\Delta H° = -74.6\ kJ\ mole^{-1}$

Example 6.8

Calculate ΔH of reaction for the following reaction, which take place when gasoline burns in internal combustion engines. Where the values of ΔH_f are –269KJ, 0 kJ, –393.5 kJ and –285 kJ for $C_8H_{18(\ell)}$, $O_{2(g))}$ $CO_{2(g)}$ and $H_2O_{(\ell)}$ respectively

$2C_8H_{18(\ell)} + 25O_{2(g)} \longrightarrow 16CO_{2(g)} + 18H_2O_{(\ell)}$

Solution:

$\Delta H°_{reaction} = \sum coeff_p \Delta H_f$ (products) $- \sum coeff_r \Delta H_f$ (reactants)

$= [16 \times \Delta H_f$ for $CO_{2(g)} + 18 \times \Delta H_f$ for $H_2O_{(\ell)}]$

$= [2 \times \Delta H_f$ for $C_8H_{18(\ell)} + 25 \times \Delta H_f$ for $O_{2(g)}]$

$= 16(-393.5\ kJ) + 18\ (-285.8\ kJ) - 2\ (-269\ kJ) - 25(0)$

$= -6296\ kJ - 5144.4\ kJ + 538\ kJ + 0$

$= -11440.4\ kJ + 538\ kJ$

$= -10902.4\ kJ$

$= -1.09 \times 10^4 kJ$

Born Haber Cycle:

Q.Explain Born-Haber Cycle by using NaCℓ as example for lattice energy calculation.

It is an application of Hess's law. It is defined as;

"Energy Change in a cycle process is always zero."

It is used to find the lattice energy of binary ionic compound.

Lattice Energy:

It is define as;

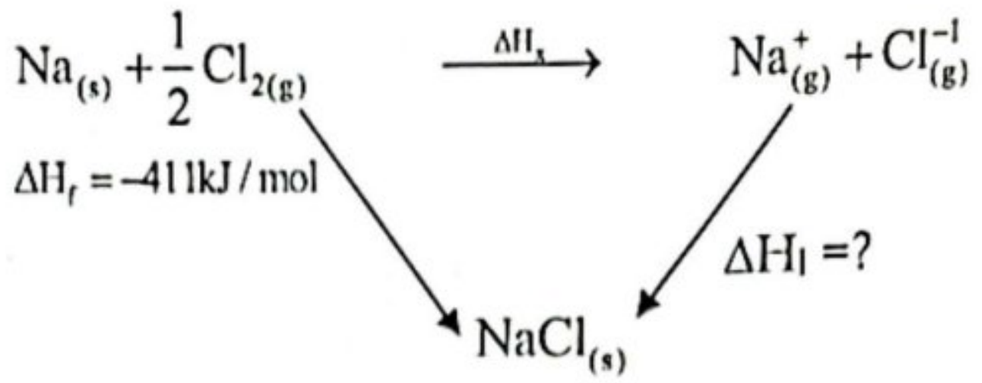
"Amount of energy released when one mole ionic compound is formed from its opposite charged ions in gaseous state."

e.g.

$Na^+_{(g)} + Cl^-_{(g)} \longrightarrow -NaCl_{(s)} \quad \Delta H_r = -787 kJ/mol$

Lattice energy depends upon size of cation and Anion. As the size of cation or anion increases, lattice energy decreases.

Energy Triangle of NaCℓ:



Step-1 $Na(s) \longrightarrow Na(g)$ $\Delta H^o_s = +108$ KJ/mole

Step-2 $Na(g) \longrightarrow Na^+(g) + e^-$ $\Delta H^o_{I.E} = +496$ KJ/mole

Step-3 $\frac{1}{2}Cl_2(g) \longrightarrow Cl(g)$ $\Delta H^o_{at} = +121$ KJ/mole

Step-4 $Cl(g) + e^- \longrightarrow Cl^-(g)$ $\Delta H^o_{E.A} = -349$ KJ/mole

$$\Delta H^o_x = \Delta H^o_{at} + \Delta H^o_{I.E} + \Delta H^o_{at} + \Delta H^o_{E.A}$$

$$= +108KJ + 496KJ + 121KJ + (-349KJ)$$

$$\Delta H^o_x = 376 kJ/mol$$

According To Hess's law;

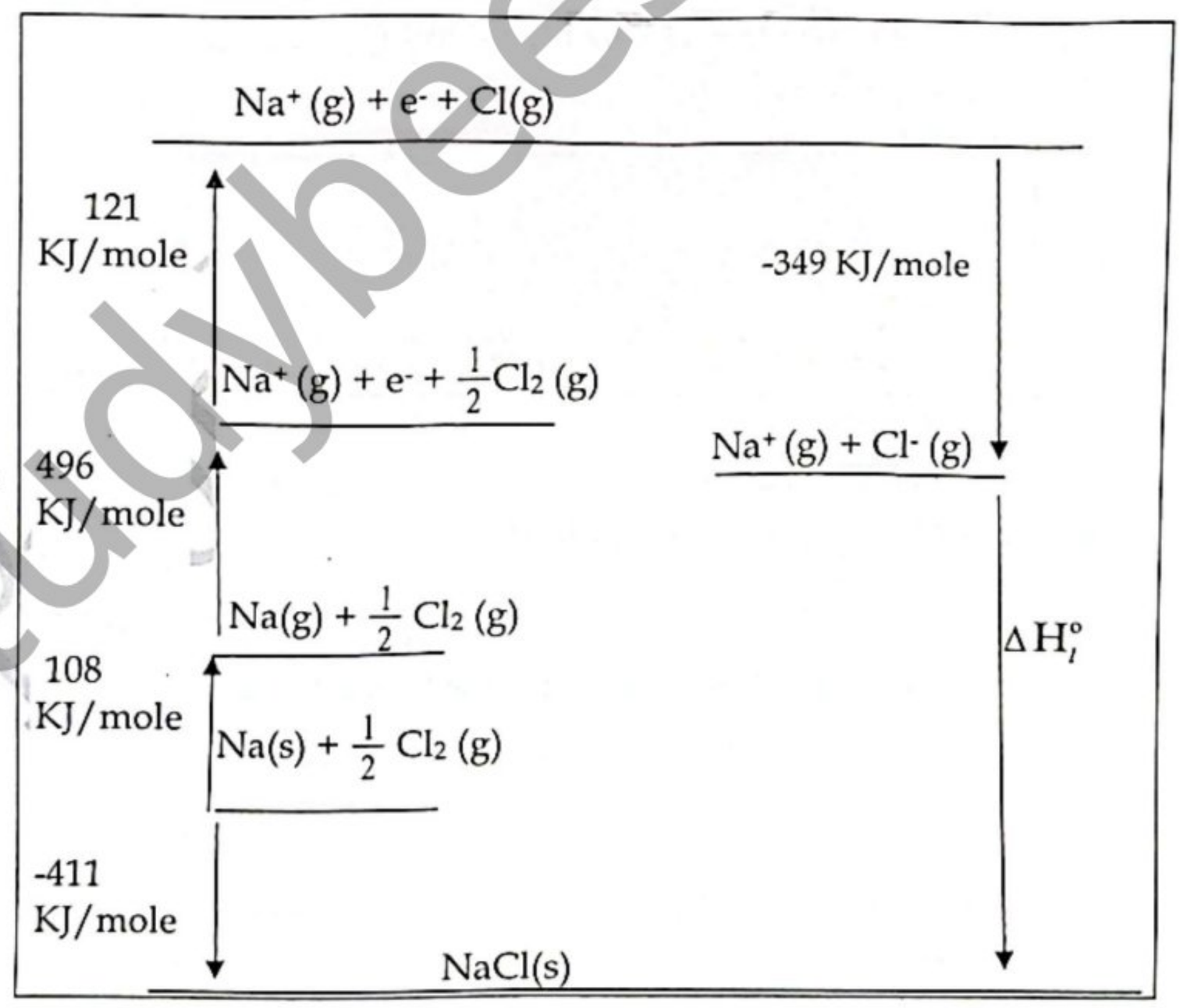
$$\Delta H_f = \Delta H_x + \Delta H_{(l)}$$

$$\Delta H_{(l)} = \Delta H_{f(l)} - \Delta H_x$$

$$= -411 - 376$$

$$\Delta H_{(l)} = -787 kJ/mol$$

Energy Cycle of NaCl:



Concept Assessment Exercise 6.5

Draw a complete Born Haber Cycle for the formation of MgO (s). Calculate the lattice energy of MgO from the following data.

1.Standard enthalpy of formation of MgO = -602 kJ $mole^{-1}$

2. Standard enthalpy of sublimation of Mg = 150 kJ $mole^{-1}$

3. Ionization energy of Mg(g) to form $Mg^{2+}(g)$ = 2180 kJ $mole^{-1}$

4. Standard enthalpy of atomization of O_2 = 247 kJ $mole^{-1}$

5.Electron affinity of O(g) to form $O^{1-}(g)$ = -141 kJ $mole^{-1}$

6. Electron affinity of $O^{1-}(g)$ to form $O^{2-}(g)$ = 878 kJ $mole^{-1}$

Steps for Lattice Energy of MgO

1. $Mg(s) + \frac{1}{2}O_2(g) \rightarrow MgO(s)$ $\quad \Delta H_f$ = -602 kJ $mole^{-1}$
2. $Mg(s) \rightarrow Mg(g)$ $\quad \Delta Hs$ = 150 kJ $mole^{-1}$
3. $Mg(g) \rightarrow Mg^{2+}(g) + 2e^-$ $\quad \Delta H_{I.E}$ = 2180 kJ $mole^{-1}$
4. $\frac{1}{2}O_2(g) \rightarrow O(g)$ $\quad \Delta Hat$ = 247 kJ $mole^{-1}$
5. $O(g) + e^- \rightarrow O^-(g)$ $\quad \Delta H_{E.A1}$ = -141 kJ $mole^{-1}$
6. $O^-(g) + e^- \rightarrow O^{2-}(g)$ $\quad \Delta H_{E.A2}$ = 878 kJ $mole^{-1}$
7. $Mg^{2+}(g) + O^{2-}(g) \rightarrow MgO(s)$ $\quad \Delta H_l$ = ?

Calculation

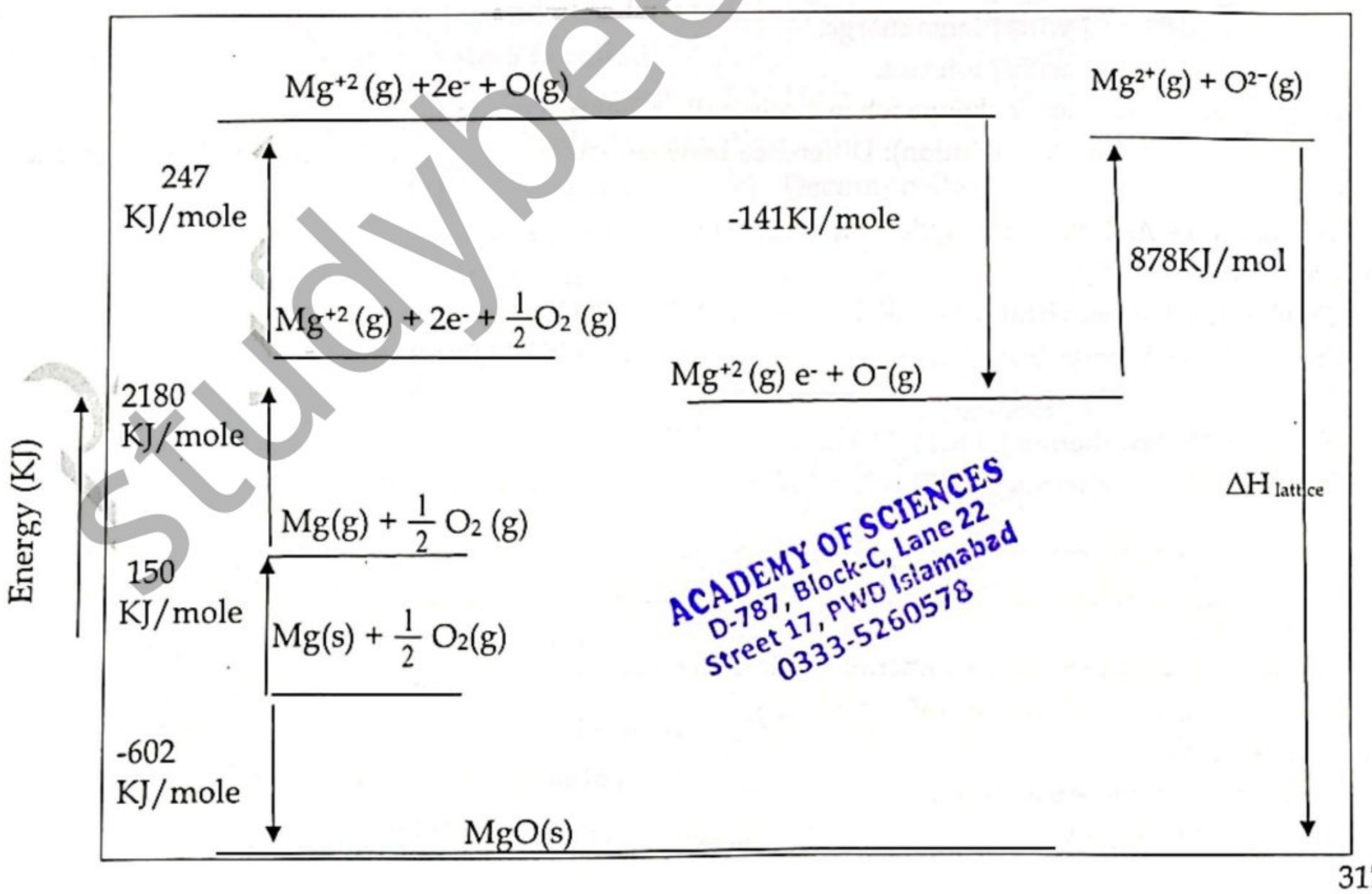
$\Delta H_f = \Delta Hs + \Delta H_{I.E} + \Delta Hat + \Delta H_{E.A1} + \Delta H_{E.A2} + \Delta H_{lattice}$

-602 kJ $mole^{-1}$ = 150 kJ $mole^{-1}$ + 2180 kJ $mole^{-1}$ + 247 kJ $mole^{-1}$ + (-141 kJ $mole^{-1}$) + 878 kJ $mole^{-1}$ + $\Delta H_{lattice}$

$\Delta H_{lattice}$ = -602 - 3314

$\Delta H_{lattice}$ = - 3916 kJ $mole^{-1}$

Born Haber cycle of MgO



Factors Affecting Lattice energy

Lattice energy depend upon charge density higher will be charge density higher will be lattice energ ioniy and vice versa

Charge density = Ionic charge/ionic radius

Two main factors influence lattice energy:

1. **Ionic Charge:** Higher charges = stronger electrostatic attraction = higher lattice energy.
2. **Ionic Radius:** Smaller radius = stronger electrostatic attraction = higher lattice energy.

Concept Assessment Exercise 6.6

Which of following has higher lattice energy and why

a. LiF vs NaF: LiF has higher lattice energy

Reason. smaller Li^+ ion

b. NaCl vs $MgCl_2$: $MgCl_2$ has higher lattice energy

Reason Mg^{2+} has higher charge than Na^+

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Role of Hydration in in the Dissolving Process

Heat of Solution: Heat released/absorbed when a solute dissolves in a solvent.

Heat of Hydration: Heat released/absorbed when a solute dissolves in water.

Main Points about Heat of Hydration

1. What happens when a solute dissolves in water:
   Forces between solute particles need to be broken.
   Solute particles bond with water molecules.
2. **Energy changes**:
   Lattice energy: Energy to break solute particles apart.
   Heat of hydration: Energy released when solute particles bond with water.
3. **Trends**:
   Heat of hydration ↑ with ↑ ionic charge.
   Heat of hydration ↓ with ↑ ion size.
4. **Heat of solvation**: Energy change when a solute dissolves in a solvent.
5. Enthalpy of solution (ΔHsolution): Difference between lattice energy (ΔHlattice) and hydration energy (ΔHhyd).
   ΔHsolution = ΔHlattice - ΔHhydration
6. Process can be:
   Exothermic: Releases heat (like NaOH in water, ΔH = -44.5 kJ/mol).
   Endothermic: Absorbs heat (like NH_4NO_3 in water, ΔH = +25.7 kJ/mol).
   Examples
   Negative ΔH (exothermic): LiCl (-37.0 kJ/mol), $AlCl_3$ (-321.0 kJ/mol).
   Positive ΔH (endothermic): NaCl (+2.98 kJ/mol), NH_4NO_3 (+25.7 kJ/mol).

Factors Affecting Enthalpy of Hydration

1. Ionic Charge

Higher ionic charge = more exothermic hydration enthalpy.

Example: MgO (Mg^{2+}, O^{2-}) > NaCl (Na^+, Cl^-).

2. Ionic Radius

Smaller ions = more exothermic hydration enthalpy.

Example: $MgSO_4$ > $CaSO_4$ because Mg^{2+} is smaller than Ca^{2+}.

Concept Assessment Exercise 6.8

LiF vs NaF: LiF has higher hydration enthalpy because Li^+ is smaller than Na^+.

$NaNO_3$ vs KNO_3: $NaNO_3$ has higher hydration enthalpy because Na^+ is smaller than K^+.

ENTROPY

What is entropy?: Measure of disorder/randomness in a system (denoted by "S").

Entropy changes:

System goes from ordered → disordered = entropy increases ($\Delta S > 0$).

Example: Ice melting.

System goes from disordered → ordered = entropy decreases ($\Delta S < 0$).

Example: Water vapor condensing.

Calculating entropy change: $\Delta S° = \Sigma S°(products) - \Sigma S°(reactants$

EXERCISE

Q1. Choose the correct Answer:

(i) Which of the following substances have zero value for their standard enthalpy of formation?

(a) O_3 (b) H_2O (c) ZnO, (d) None of these

(ii) Calorie is equivalent to

(a) 4.18J (b) 4.18 kJ (c) 0.418J (d) 0.418 kJ.

(iii) Enthalpy of neutralization of all the strong acids and strong basis has the same value due to

(a) The formation of salt and water (b) The formation of salt

(c) The complete ionization of acids and bases

(d) The combination of H^+ and OH^- ions to form water.

(iv) Total heat content of a system is called

(a) Enthalpy (b) Internal energy (c) Heat (d) State function

(v) Heat of _____ of a substance is always negative

(a) Formation (b) Combustion (c) Decomposition (d) Solution

(vi) A balloon filled with oxygen is placed in a freezer. Identify system

(a) Balloon (b) Oxygen (c) Freezer (d) All of these

(vii) A bomb calorimeter is used in _____ calorimetry:

(a) Constant volume (b) Constant pressure

(c) Both a and b (d) Constant temperature

(viii) Boron Haber cycle is used to determine lattice energies of

(a) Molecular solids (b) Ionic solids (c) Covalent solids (d) Metallic solids

(ix) Enthalpy of combustion for C is $-393.5\ kJmole^{-1}$

$C(s) + O_2(g) \rightarrow CO_2(g)$ $\Delta H_{Combustion} = -393.5 kJmole^{1}$

Enthalpy of formation of CO_2 would be

(a) +393.5 kJ (b) –393:5 kJ

(c) Zero (d) Cannot be predicted form the given equation.

(x) Which of the following is not a state function of a system?

(a) Thermal energy at constant pressure (b) Enthalpy

(c) Internal energy (d) Work done

(xi) **For writing a thermochemical equation for enthalpy of combustion of an element requires,**

(a) 1 mole of element as reactant

(b) 1 mole of oxide of element as product

(c) Standard states of all the substances

(d) Balanced equation of 1 mole of element

1. (a) 2. (a), (b) 3. (a), (c), (d) 4. (a), (b), (c), (d)

======

ANSWERS

i. (d)	ii. (a)	iii. (c)	iv. (a)	v. (b)	vi. (b)	vii. (a)	viii. (b)	ix. (b)
x. (d)	xi.3a,c,d							

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SHORT QUESTIONS/ANSWERS

Q.2 Name and define units of thermal energy

Ans. Units of thermal energy

Unit of heat or thermal energy used in SI system is the Joule (J).

It is defined as the energy required moving an object or a particle through a distance of one meter by a force of one Newton.

Another common unit of heat is the calorie. It is defined as the heat or thermal energy required to raise the temperature of one gram of water from 14.5 to 15.5°C.

1 Calorie = 4.18 Joules

Q.3 Define the following terms system, surroundings and boundary.

Ans:

- **System:** The part of the universe on which we wish to focus attention is called a system.
- **Surroundings:** The part of the universe except system is called surroundings.
- **Boundary:** The real or imaginary surface separating the system from surroundings is called boundary.
- **STATE FUNCTION:** A property of a system that is determined by the state of the system regardless of how that condition was achieved is called state function.
- **HEAT:** Heat is the transfer of thermal energy between two bodies that are at different temperature.
- **HEAT CAPACITY:** The amount of heat required to raise the temperature of the given quantity of a substance by one Kelvin is called heat capacity.

  INTERNAL ENERGY: The sum of all kinds of energies of all the particles of a system is called internal energy.

  i.e. K = K.E + P.E
- **Enthalpy of substance:**

  Sum of internal energy and product of pressure, valour is known as enthalpy.

Q.4 Classify reactions as exothermic or endothermic.

Ans. See notes

Q.5 Define bond dissociation energy.

Ans. The amount of energy required to break one mole of a particular bond to form neutral atoms is called bond dissociation energy.

Q.6 When ethanol burns in oxygen, carbon dioxide and water are formed

(a) Write the equation which describes this reaction.

(b) Using the following data, calculate the enthalpy of combustion for ethanol C_2H_5OH.

ΔH_f^o for ethanol(l) = -277.0 KJmole^{-1}

ΔH_f° $CO_2(g)$ = -393.5 KJ mole^{-1}, ΔH_f° water(l) = -285.8KJmole^{-1}

Ans.

eq (1) $2C_{(s)} + 3H_{2(g)} + {}^1/_2 O_{2(g)} \rightarrow C_2H_5OH_{(1)}$ ΔH_f° =-277.0 kJ mole^{-1}

eq (2) $C(s) + O_{2(g)} \longrightarrow CO_{2(g)}$ ΔH_f° =–393.5 kJ mole^{-1}

eq (3) $H_{2(g)} + \frac{1}{2}O_{2(g)} \longrightarrow H_2O_{(1)}$ ΔH_f° =–285.8 kJ mole^{-1}

Multiplying eq (2) with '2'

$2C_{(s)} + 2O_{2(g)} \longrightarrow 2CO_{2(g)}$ ΔH =-787.0 kJ

Multiplying eq (3) with '3'

$3H_{2(g)} + \frac{3}{2}O_{2(g)} \longrightarrow 3H_2O_{(1)}$ ΔH =-857.4 kJ

Adding eq (4) & (5)

$2C + 3H_2 + \frac{7}{2} O_2 \longrightarrow 2CO_2 + 3H_2O$ ΔH =-1644.4 kJ mol^{-1}

eq (1) $C_2H_5OH_{(1)} \longrightarrow 2C_{(s)} + 3H_{2(g)} + \frac{1}{2}O_{2(g)}$ ΔH =+277.0 kJ

(subtracting)

$C_2H_5OH_{(1)} + 3H_{2(g)} \longrightarrow 2CO_{2(g)} + 3H_2O_{(1)}$ ΔH_c° = -1367.4 kJ mole^{-1}

Q.7 Calculate from the data in table 11.1 and 11.2 the enthalpy changes of the following reactions at standard conditions.

(a) $C_6H_6(l) \longrightarrow 6C(s) + 3H_2(g)$

Ans.

eq (1) $C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)}$ ΔH° = -393.5 kJ mole^{-1}

eq (2) $H_{2(g)} + \frac{1}{2}O_{2(g)} \longrightarrow H_2O_{(1)}$ ΔH° = -285.8 kJ mole^{-1}

eq (3) $C_6H_{6(1)} + \frac{15}{2} O_{2(g)} \longrightarrow 6CO_{2(g)} + 3H_2O_{(1)}$ ΔH° = -3268 kJ mole^{-1}

Multiplying eq (1) with '6'

$6CO_{2(g)} \longrightarrow 6C_{(s)} + 6O_{2(g)}$ ΔH° = +2361.0 kJ^{-1}

Multiply Eq. (3) by 3 and reverse the equation,

$3H_2O_{(l)} + \frac{3}{2}O_{2(g)} \longrightarrow 3H_{2(g)}$ ΔH° = +857.4 kJ mol^{-1}

Adding eq (4) & (5)

eq (3) $6CO_2 + 3H_2O_{(1)} + \frac{15}{2} O_{2(g)} \longrightarrow 6CO_{2(g)} + 3H_2O_{(1)}$ ΔH° = -3218 kJ mole^{-1}

(subtracting)

$6CO_2 + 3H_2O_{(1)} \longrightarrow C_6H_{6(l)} + \frac{15}{2} O_{2(g)}$ ΔH° = -3268 kJ mole^{-1}

$6C_{(s)} + 3H_{2(g)} \longrightarrow C_6H_{6(1)}$ ΔH° = -49.6 kJ mol^{-1}

Q.8 The heat of combustion of liquid benzene, C_6H_6 to form H_2O and CO_2 at 1 atm and 25°C is –3268 KJmole^{-1} of benzene. What is the heat of formation of liquid benzene under these conditions?

Ans. Heat of combustion of liquid benzene $C_6H_6(l)$

$C_6H_{6(l)} + \frac{15}{2} O_{2(g)} \rightarrow 6CO_{2(g)} + H_2O_{(l)}$ $\quad \Delta H^\circ = -3286$ KJ/mole -----1

Heat of formation of gaseous $CO_{2\,(g)}$

$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$ $\quad \Delta H^\circ = -393.5$ KJ/mole ----2

Heat of formation of liquid $H_2O_{(l)}$

$H_{2(g)} + \frac{1}{2} O_{2(g)} \rightarrow H_2O_{(l)}$ $\quad \Delta H^\circ = -285.8$ KJ/mole ---- 3

Heat of formation of $C_6H_{6(l)}$

$6C_{(s)} + 3H_{2(g)} \rightarrow C_6H_{6(l)}$ $\quad \Delta H_f^\circ = ?$

Multiply eq.2 by 6

Multiply eq.3 by 3 and reverse eq.1

Then add all the resulting equation, we get

$6C_{(s)} + 6O_{2(g)} \rightarrow 6CO_{2(g)}$ $\quad \Delta H^\circ = -2361.0$ KJ/mole

$3H_{2(g)} + \frac{3}{2} O_{2(g)} \rightarrow 3H_2O_{(l)}$ $\quad \Delta H^\circ = -857.4$ KJ/mole

$6CO_{2(g)} + H_2O_{(l)} \rightarrow C_6H_{6(l)} + \frac{15}{2} O_{2(g)}$ $\quad \Delta H^\circ = +3286$ KJ/mole

$6C_{(s)} + 3H_{2(g)} \rightarrow C_6H_{6(l)}$ $\quad \Delta H_f^\circ = +49.6$KJ/mole

Q.10 Camphor ($C_{10}H_{16}O$) has a heat of combustion of 5903.6 KJ/mole. A sample of camphor having mass of 0.1204g is burned in a bomb calorimeter. The temperature increases by 2.28°C. Calculate heat capacity of the calorimeter.

Ans. Mass of camphor m= 0.1204g

Temperature $\Delta T = 2.28\ ^\circ C$ OR 2.28K

Enthalpy of combustion = $\Delta H_c^\circ = 5903.6$ KJ /mole

Molar mass of camphor ($C_{10}H_{16}O$) = 120+16 +16 =152 g/mole

Number of moles of camphor $= \frac{0.1204}{152} = 7.92 \times 10^{-4}$ moles

1 mole of camphor produces heat = 5903.6 KJ

7.92×10^{-4} moles of camphor produce heat = $5903.6 \times 7.92 \times 10^{-4}$

= 4.675 KJ

In a bomb calorimeter, heat evolved = $\Delta T \times$ Heat capacity of calorimeter

$4.675 = 2.28 \times$ Heat capacity of calorimeter

$$Heat\ capacity\ of\ calorimeter = \frac{4.675}{2.28} = 2.05\ KJK^{-1}$$

Q.11 In a coffee-cup calorimeter 100cm³ of 1.0M HCl and 100 cm³ of 1.0M NaOH are mixed at 24.6°C raised temperature by 6.9°C. Calculate the enthalpy of neutralization of HCl by NaOH from the given data. Heat capacity of water is 4.18 J/g°C.

Ans. Volume of NaOH = 100cm³

Volume of HCl = 100cm³

Total volume of reaction mixture = 100cm³ +100cm³ =200g

Density of water = 1g cm^{-3}

Total mass of reaction mixture (m) = 200cm3 × 1g cm-3 =200g

Rise in temperature ΔT = 6.9°C

Heat capacity of water =C = 4.18 $Jg^{-1}{}^{o}C^{-1}$

Now

$q = m \times C \times \Delta T$

$q = 200g \times 4.2J\ g^{-1}{}^{o}C^{-1} \times 6.9{}^{o}C$

$q = 5.796 \times 10\ 3J$

$$\text{No. Of moles of NaOH} = \frac{\mathbf{M \times Vol\ of\ solution\ (cm3)}}{\mathbf{1000}}$$

$$= \frac{\mathbf{1 \times 100}}{\mathbf{1000}} = 0.10 \text{ moles}$$

$$\text{Similarly, No of moles of HCl} = \frac{\mathbf{1 \times 100}}{\mathbf{1000}} = 0.10 \text{ moles}$$

Considering chemical equation

$NaOH\ (aq) + HCl(aq) \rightarrow NaCl\ (aq) + H_2O(l)$

1 mole 1 mole 1 mole

0.1 mole 0.1 mole 0.1 mole

Thus formation of 0.1 moles of water release heat = 5.796 × 10 3J

$$\text{Formation of 1 mole of water will release heat} = \frac{5.796 \times 10_{3}}{0.1_{0}}$$

= 5.796 × 10 4J mole-1

=57.96 kJ mole -1

Since heat is evolved at constant pressure,

$q_p = \Delta H_n^o = -57.96\ kJ\ mole\ -1$

Q.12 Calculate ΔH° for the reaction

$S(s) + O_2(g) \rightarrow SO_2(g)$

from the following data

$S(s) + \frac{3}{2} O_2(g) \rightarrow SO_3(g) \quad \Delta H^\circ = -395.2$ KJ

$2SO_2(g)\ O_2(g) \rightarrow 2SO_3(g) \quad \Delta H^\circ = -198.2$ KJ

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Ans. Heat of combustion of liquid benzene C_6H_6 (1)

$S(s) + \frac{3}{2} O_2(g) \rightarrow SO_3(g) \quad \Delta H^o = -395.2$ KJ _____ (1)

$2SO_2(g)\ O_2(g) \rightarrow 2SO_3(g) \quad \Delta H^o = -198.2$ KJ _____ (2)

Heat of formation of C_6H_6 (1)

$S(s) + O_2(g) \rightarrow SO_2(g) \quad \Delta H^o = ?$

Divide Eq. (2) by 2,

Take Eq. As such

Then add all the resulting equation. We get

$S(s) + \frac{3}{2} O_2(g) \rightarrow SO_3(g) \quad \Delta H^o = -395.2$ KJ _____ (1)

$SO_3(g) \rightarrow SO_2(g) + \frac{1}{2} O_2(g) \quad \Delta H^o = +99.1$ KJ _____ (2)

$S(s) + O_2(g) \rightarrow SO_2(g) \quad \Delta H^o = +\ -296.1$ KJ

Q.13 An aluminum frying pan weighs 745g is heated on a stove from 25°C to 205°C. What is q for the frying pan? C_P for Al is 24.35 $Jmole^{-1}K^{-1}$

Ans. Mass of aluminum frying pan = 745 g

Atomic mass of aluminum frying pan =27 g mole -1

Number of moles of aluminum = $n = \frac{745}{27}$ = 27.596 moles

Initial temperature = T_1 = 25°C

Final temperature = T_2 = 205°C

Temperature rise = ΔT 205 °C –25 °C

= 180 °C

At constant pressure heat capacity of Al = Cp = 24.35 J mole-1 K-1

At constant pressure , the heat released can be determined as

$q = n \times Cp \times \Delta T$

= 27.596 mole × 24.35 J mole -1 K^{-1} × 180K

= 120953J

=120.953 kJ

Q.14 Write the balanced equation for the formation reaction of each of the following substances

(i) C_4H_9OH (Butanol) (ii) Rust, Fe_3O_4

(iii) Acetic acid CH_3CO_2H (iv) Urea $(NH_2)_2CO$

Ans. (i) $4C_{(s)} + 5H_{2(g)} + \frac{1}{2}O_{2(g)} \rightarrow C_4H_9OH_{(l)}$

(ii) $3Fe(s) + 2O_2(g) \rightarrow Fe_3O_4(s)$

(iii) Acetic acid CH_3CO_2H

$2C_{(s)} + 2H_{2\,(g)} + O_{2\,(g)} \rightarrow CH_3COOH_{(l)}$

(iv) Urea $(NH_2)_2CO$

$C_{(s)} + N_{2\,(g)} + 2H_{2\,(g)} + \frac{1}{2}O_{2\,(g)} \rightarrow (NH_2)_2CO_{(s)}$

Q.15 The human body burns glucose for energy. Burning 1.0g of glucose produces 15.65 KJ of heat.

(a) Write the balanced equation for the combustion of glucose.

(b) Determine the molar heat of combustion of glucose.

Ans.

(a) $C_6H_{12}O_{6(s)} + 6O_{2(g)} \rightarrow 6CO_{2(g)} + 6H_2O_{(l)}$ ΔH_c^o = -2817.28 kJ $mole^{-1}$

(b) Give mass of glucose = 1.0 g

Molar mass of camphor ($C_6H_{12}O_6$) = 12 × 6 + 1 × 12 + 16 × 6

= 180 g $mole^{-1}$

Number of moles of glucose = $\frac{1}{180}$ = 5.56 × 10^{-3} moles

Given heat produced by glucose = 15.65 kJ

5.555 × 10^{-3} moles of glucose produce heat = 15.65 kJ

1 mole of glucose produce heat = $\frac{15.65}{5.555 \times 10 - 3}$

= -2817.28 kJ

(c) Heats of combustions of C and H_2 are -393.5KJmole^{-1}and -285.8KJmolre^{-1} respectively. Determine the heat of formation of glucose.

Ans. Heat of combustion of glucose

$C_6H_{12}O_{6(s)} + 6O_{2(g)} \rightarrow 6CO_{2(g)} + 6H_2O_{(l)}$ $\Delta H = -2817.28$ kJ mole^{-1} —— 1

Heat of combustion of C

$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$ $\Delta H^o = -393.5$kJ mole^{-1} —— 2

Heat of combustion of H_2

$H_{2(g)} + \frac{1}{2}O_{2(g)} \rightarrow H_2O_{(l)}$ $\Delta H^o = -285.8$ kJ mole^{-1} —— 3

Heat of combustion of $C_6H_{12}O_{6(s)}$

$6C_{(s)} + 6H_{2(g)} + 3O_{2(g)} \rightarrow C_6H_{12}O_{6(s)}$ ΔH^o =?

Multiply Eq. (2) by 6,

Multiply Eq. (3) by 6,

Reverse Eq. (1)

Then add all the resulting equations. We get

$6C_{(s)} + 6O_{2(g)} \rightarrow 6CO_{2(g)}$ $\Delta H^o = -2361$ kJ

$6H_{2(g)} + 3O_{2(g)} \rightarrow 6H2O_{(l)}$ $\Delta H^o = -1714.8$ kJ

$6CO_{2(g)} + 6H_2O_{(l)} \rightarrow C_6H_{12}O_{6(s)} + 6O_{2(g)}$ $\Delta H^o = +2817.28$ kJ

$6C_{(s)} + 6H_{2(g)} + 3O_{2(g)} \rightarrow C_6H_{12}O_{6(s)}$ $\Delta H^o = -1258.52$ kJ mole^{-1}

Q.16 The standard combustion enthalpies of carbon, hydrogen and acetic acid are -393.5 KJ mole$^-$, -285.8KJmole^{-1} and –875 KJmole^{-1} respectively. Deduce the value of standard enthalpy of formation of acetic acid, CH_3COOH.

Ans. Heat of combustion of glucose

$CH_3COOH_{(l)} + 2O_2 \rightarrow 2CO_{2(g)} + 2H_2O_{(l)}$ $\Delta H^o = -875$ kJ mole^{-1} —— 1

Heat of combustion of C.

$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$ $\Delta H^o = -393.5$ kJ mole^{-1} —— 2

Heat of combustion of H_2.

$H_{2(g)} + \frac{1}{2}O_{2(g)} \rightarrow H_2O_{(l)}$ $\Delta H^o = -285.8$ kJ mole^{-1} — 3

Heat of combustion of liquid $C_2H_3OH_{(l)}$

$2C_{(s)} + 2H_{2(g)} + O_{2(g)} \rightarrow CH_3COOH_{(l)}$ ΔH^o = ?

Multiply Eq. (2) by 2

Multiply Eq. (3) by 2,

Multiply Eq. (1)

Then add all the resulting equations, we get.

$2C_{(s)} + 2O_{2(g)} \rightarrow 2CO_{2(g)}$ $\Delta H^o = -787$ kJ

$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(l)}$ $\Delta H^o = -571.6$ kJ

$2CO_{2(g)} + 2H_2O_{(l)} \rightarrow CH_3COOH_{(l)} + 2O_{2(g)}$ $\Delta H^o = +875$ kJ

$2C_{(s)} + 2H_{2(g)} + O_{2(g)} \rightarrow CH_3COOH_{(l)}$ $\Delta H^o = -483.6$ kJ mole^{-1}

Q.17 Is the conversion of magnetite, Fe_3O_4 to hematite, Fe_2O_3 by oxygen is endothermic or exothermic? Justify your answer.

$3Fe(s) + 2O_2(g) \rightarrow Fe_3O_4(s)$ $\Delta H^\circ = -1118$ KJ

$2Fe(s) + \frac{3}{2} O_2(g) \rightarrow Fe_2O_3(s)$ $\Delta H^\circ = -824.2$ KJ

Ans.

$3Fe(s) + 2O_2(g) \rightarrow Fe_3O_4(s)$ $\Delta H = -1118$ kJ ___ (1)

$2Fe(s) + \frac{3}{2} O_2(g) \rightarrow Fe_2O_3(s)$ $\Delta H + -824.2$ kJ ____ (2)

The required conversion is

$2Fe_3O_4(s) + O2 \rightarrow 3Fe_2O_3(s)$ $\Delta H = ?$

The standard enthalpy of formation of an element its pure form is taken as zero.

Thus, according to the relation

$\Delta H^o_{reaction} = \sum \text{coeff}_p \Delta H^o_f$ (products) $-\sum$ coeffr ΔH^o_f (reactants)

$= [3\times (-824.2)] - [2\times (-1118) + (0)]$

$= (-2472.6) - (-2236) = -2472.6 + 2236 = -236.6$ kJ

ADDITIONAL MCQS

IMPORTANT MCQ's

1. **If an endothermic reaction is allowed to take place very rapidly in the air. The temperature of the surrounding air.**
   (a) remains constant (b) increases
   (c) decreases (d) remains unchanged
2. **In endothermic reactions, the heat contents of the**
   (a) products is more than that of reactants.
   (b) reactants is more than that of products.
   (c) both (a) and (b) (d) products is equal to reactants
3. **A chemical change always involve**
   (a) absorption of heat (b) emission of heat
   (c) either absorption or emission of heat
   (d) the liberation of heat and light energy
4. **Spontaneous processes are mostly**
   (a) reversible (b) irreversible (c) endothermic (d) none of these
5. **Which is not a state function**
   (a) enthalpy (b) entropy (c) pressure (d) heat
6. **Which of the following statements is contrary to the first law of thermodynamics?**
   (a) Energy can neither be created nor destroyed.
   (b) One form of energy can be transferred into an equivalent amount of other kind of energy.
   (c) In an adiabatic process, the work done is independent of its path.
   (d) Continuous production of mechanical work without supplying equivalent amount of heat is possible.
7. **For a given process, the heat changes at constant pressure (q_p) and at constant volume (q_v) are related to**
   (a) $q_p = q_v$ (b) $q_p < q_v$ (c) $q_p > q_v$ (d) $q_p = q_v/2$
8. **First law of thermodynamics is represented as**
   (a) $\Delta E = q + RT$ (b) $\Delta E = \Delta H$ (c) $\Delta E = q + \Delta p$ (d) $\Delta E = q + w$
9. **The net heat change in a chemical reaction is same whether it is brought about in one or several steps. It is known as**

(a) Henry's law (b) Joule's principle
(c) Hess's law (d) Law of conservation of energy

10. **Hess' s law of heat summation includes**
(a) initial reactants only (b) initial reactants and final products
(c) final products only (d) intermediates only

11. **Enthalpy change of which compound cannot be measured directly**
(a) CO_2 (b) NO_2 (c) SO_2 (d) CO

12. **Born Haber cycle is used to determine the**
(a) lattice energy (b) enthalpy of sublimation
(c) enthalpy of vaporization (d) enthalpy of neutralization

13. **The smallest Unit of heat energy is**
(a) calorie (b) joule (c) erg (d) kilowatt hour

14. **Most of the reaction with more stable product are**
(a) exothermic (b) endothermic
(c) reaction (d) reaction ionic reaction

15. **For the reaction $NaOH + HC\ell \longrightarrow NaCl + H_2O$ the change in enthalpy is called.**
(a) heat of reaction (b) heat of formation
(c) heat of neutralization (d) heat of combustion

16. **The change in heat energy of a chemical reaction at constant temperature and pressure is called.**
(a) enthalpy change (b) heat of sublimation
(c) bond energy (d) internal energy change

17. **One Calorie is equivalent to**
(a) 0.418 J (b) 41.84J (c) 4.184J (d) 418.4J

18. **Enthalpy of neutralization of .all the strong acids and strong bases has the same value because**
(a) Neutralization leads to the formation of salt and water.
(b) Strong acids and strong bases are ionic substances
(c) Acids always give H^+ ions and bases always furnish OH^- ions.
(d) The net chemical change involve the combination of H^+ and OH^- ions to form water

19. **Bomb calorimeter cannot measure the enthalpy of formation of:**
(a) CO_2 (b) SO_2 (c) MgO (d) All of the above

20. **In an endothermic reaction ΔH is**
(a) positive (b) negative (c) zero (d) all

21. **The enthalpy of an element in standard states is**
(a) 1 kJ mol^{-1} (b) zero (c) 298kJ mol^{-1} (d) always positive

22. **The Standard heat of formation is measured at 1 atmosphere and**
(a) 0°C (b) 100°C (c) 273°C (d) 25°C

23. **The heat contents of the system is known as**
(a) entropy (b) enthalpy (c) work (d) free energy

24. **The heat of neutralization of given reaction is -57 kJ**

$$NaOH + HC1 \rightarrow NaCl + H_2O$$

19 **By using this information, what is the value for heat liberated in the following neutralization reaction.**

$$Ba(OH)_2 + 2HC\ell \rightarrow BaC\ell_2 + 2H_2O$$

(a) -57 kJ (b) -76kJ (c) -114kJ (d) -228kJ

25. **In a bomb calorimeter, the reactions are carried out at**

(a) constant volume (b) constant pressure
(c) constant temperature (d) a, b and c conditions

26. The enthalpy of combustion is

(a) positive (b) negative
(c) either positive or negative (d) no correlation

27. ΔH for the neutralization reaction NaOH and CH_3COOH is

(a) -57kJ (b) less than -57 kJ (c) greater than -57 kJ (d) zero

28. Which one of the following enthalpies is always an exothermic process

(a) ΔHc (b) ΔHs (c) ΔH_{at} (d) ΔH_f

Answer

1.		2.		3.		4.		5.		6.		7.		8		9.		10.	
11.		12.		13.		14.		15.		16.		17.		1		19.		20.	
21.		22.		23.		24.		25.		26.		27.		2					