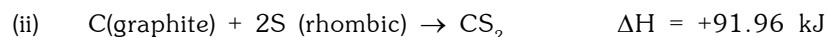
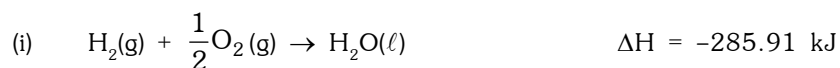


THERMOCHEMISTRY

□ THERMOCHEMISTRY :

All chemical processes are associated with energy changes in chemical reactions. A thermochemical equation, represents both the material change and the energy. While writing a thermochemical equation, the heat evolved in case of exothermic reaction, or the heat absorbed in case of endothermic reaction, is indicated on the product side of the balanced chemical equation. For example,



The most stable physical state of the reacting species and the products is also indicated in brackets. Thus, a thermochemical equation gives complete information about the material change and the associated heat change. It is clear that the first reaction is exothermic and the second one endothermic.

◆ Examples of exothermic reactions

Coal is burnt simply for the large amount of energy available during its combustion :

When water is added to quick lime (CaO) for preparing whitewash, a considerable amount of heat is produced during the reaction. The heat produced warms up the water.

When we add dilute hydrochloric acid to a test tube containing granulated zinc, hydrogen gas is evolved. The reaction is accompanied by evolution of heat.

◆ Examples of endothermic reactions.

When a small quantity of ammonium chloride (NH₄Cl) is dissolved in water in a test tube, the tube becomes colder than before. During this chemical reaction heat is absorbed from the surroundings (test tube).

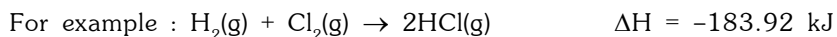
When the crystals of sodium thiosulphate (Na₂S₂O₃·5 H₂O) commonly called hypo, are dissolved in water, a cooling effect takes place.

One mole of nitrogen reacts with one mole of oxygen to form two moles of nitric oxide. 180.5kJ of heat is absorbed at constant temperature and the reaction may be expressed as:



□ STANDARD ENTHALPY OF REACTION :

Enthalpy of reaction is defined as the quantity of heat evolved or absorbed. When molar quantities of substances react in amounts represented by chemical equation at constant pressure and temperature, the substances being in their standard states.

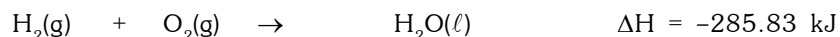


Thus, when one mole of gaseous hydrogen reacts with one mole of gaseous chlorine, 183.92 kJ of heat is evolved. This is the heat of the reaction for the above chemical reaction.

$$\Delta H_r = \sum H_{\text{products}} - \sum H_{\text{reactants}}$$

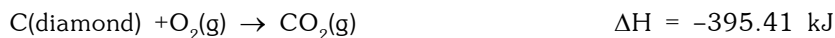
◆ Factors Affecting Heat or Enthalpy Change : The enthalpy change of reaction (ΔH) depends upon the following factors :

◆ Physical state of the reactants and the products : The ΔH of a reaction depends upon the physical states of reactants and products. For example, when hydrogen and oxygen gases combine to give liquid water, the heat of reaction is different than when they combine to form gaseous water.



- ◆ **Quantities of reactants** : The amount of heat evolved or absorbed depends upon the amount of reactants. For example, the heat of combustion of 2 moles of carbon is double than heat of combustion of 1 mole of carbon.

- ◆ **Allotropic modification** : The amount of heat evolved or absorbed for different forms of the same substance are different. For example,



- ◆ **Temperature** : The heat of reaction depends upon the temperature of reactants and products.



- ◆ **Pressure or volume**: The heat of reaction depends upon the conditions of constant pressure or volume. As,

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

ΔH may be equal, greater than or less than ΔU .

Thus, a reaction, which is exothermic in one direction will be **endothermic in reverse direction**.

- ◆ **Standard Enthalpy of Reaction** : As the heat of a reaction varies with temperature it is desirable to fix up a standard or a reference state. A substance is said to be in standard state when it is present in its most stable state at 298 K under a pressure of one atmosphere.

The heat change that occurs when a process is carried out at 298K and one atmospheric pressure is called standard heat change. It is represented as ΔH° where the superscript (°) indicates the standard state. If the reactants and products are in their standard states, then the enthalpy of a reaction is termed as the standard enthalpy of reaction

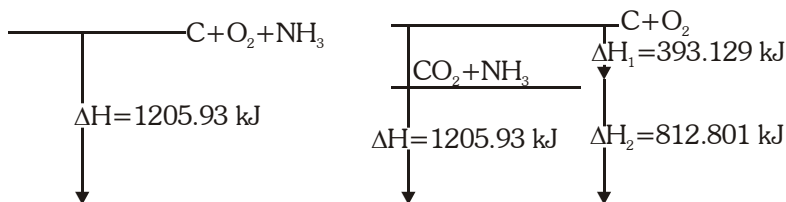
Note : The thermodynamic standard state can be established at any temperature, but most book tabulate standard enthalpy at 298.15K. The superscript (°) indicate rather pressure of 1 bar. Thus there can be one standard state at each temperature.

□ HESS'S LAW OF CONSTANT HEAT SUMMATION :

This law put forth by Hess states that "the resultant enthalpy change in a reaction is the same whether it occurs in one or several steps".

We illustrate this law by considering a example :

- The formation of urea from carbon, oxygen and ammonia may take place directly in one step or in two steps as shown below :



(i) One step reaction

(ii) Two step reaction

The enthalpy change when the reaction takes place in two steps is given by

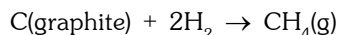
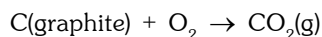
$$\Delta H_1 + \Delta H_2 = -393.129 - 812.801 = -1205.93 \text{ kJ}$$

This is seen to be equal to ΔH , the change involved in one step reaction.

◆ **Enthalpy of Formation :**

The heat evolved or absorbed when **1 mole** of a substance is formed from its constituent elements under constant pressure and temperature conditions is called enthalpy of formation. It is denoted by ΔH_f .

For example, heat of formation of carbon dioxide and methane may be expressed as :



These equations should always be written for one mole as per the definition of the substance to be formed. If for balancing, we require the coefficient 2, 3.... etc., in the equation then ΔH_f values should also be multiplied by the same number as discussed earlier.

Ex. Calculate the standard heat of formation of carbon disulphide (ℓ). Given that the standard heats of combustion of carbon (s) sulphur (s) and carbon disulphide (ℓ) are 393.3, -293.72 and -1108.76 kJ mol⁻¹ respectively.

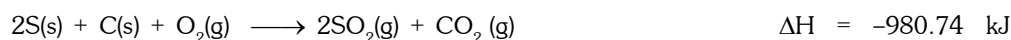
Sol. The given data can be written in thermochemical equation form as :



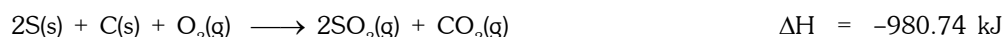
The required equation is:



Multiplying equation (ii) by 2 and adding to equation (i) we get,



Subtracting equation (iii) from the above equation we have,



◆ **Standard enthalpy of formation :** The heat of formation ΔH_f depends upon the condition of temperature, pressure and physical state (gas, liquid or solid) of the reactants and the products. Therefore the heat change accompanying the formation of one mole of a compound from its elements when all the substances are in their standard states (1 atm pressure and 298 K), is called the standard heat of formation. It is expressed as ΔH_f° .

Since no heat changes are involved in the formation of elements from themselves in their standard states, the standard enthalpy of formation of all elements is zero.

For example, the standard enthalpy of formation (ΔH_f°) for $\text{H}_2\text{O}_{(\ell)}$ is - 286 kJ mol⁻¹ i.e., when one mole of liquid water is formed from its elements $\text{H}_{2(\text{g})}$ and $\text{O}_{2(\text{g})}$ at 298 K and 1 atm pressure, then 286 kJ mol⁻¹ of heat is released. The negative value of ΔH_f° indicates the formation of a stable compound.

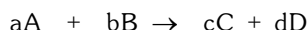
◆ **Standard heat of reaction from standard heats of formation :**

The knowledge of standard heats of formation of various substances can be used to calculate the heats of reactions under standard conditions. The standard heat of any reaction (ΔH) is equal to the difference between the ΔH_f° of all the reactants i.e.,

$\Delta H = \text{Sum of the standard heats of formation of products} - \text{Sum of the standard heats of formation of reactants.}$

$$\text{i.e., } \Delta H^\circ = \sum \Delta H_f^\circ (\text{products}) - \sum \Delta H_f^\circ (\text{reactants})$$

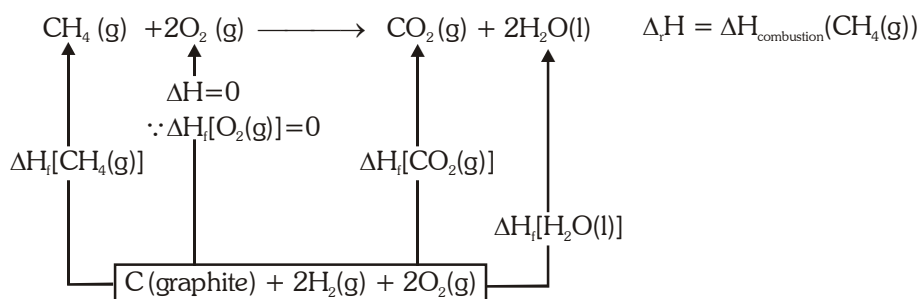
For a reaction,



$$\begin{aligned} \Delta H &= \sum \Delta H_f^\circ (\text{products}) - \sum \Delta H_f^\circ (\text{reactants}) \\ &= [c\Delta H_f^\circ (C) + d\Delta H_f^\circ (D)] - [a\Delta H_f^\circ (A) + b\Delta H_f^\circ (B)] \end{aligned}$$

Consider the combustion of CH_4 , the enthalpy change of the reaction is equal to the enthalpy of combustion of CH_4 .

$$\Delta H_r = [\Delta H_f (\text{CO}_{2(g)}) + 2\Delta H_f (\text{H}_2\text{O}_{(l)})] - [2\Delta H_f (\text{O}_{2(g)}) + \Delta H_f (\text{CH}_4)]$$

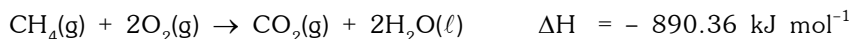


□ **ENTHALPY OF COMBUSTION :**

Enthalpy of combustion of a given compound is defined as follows :

It is the enthalpy change when one mole of this compound combines with the requisite amount of oxygen to given products in their stable forms.

For example, the standard enthalpy of combustion of methane at 298.15 K is $-890.36 \text{ kJ mol}^{-1}$. This implies the following reaction :

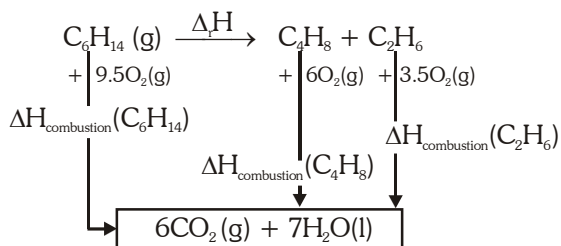


The standard enthalpy of combustion of methane at 298.15 K may be written as

$$\Delta_c H (\text{CH}_4, g, 298.15 \text{ K}) = -890.36 \text{ kJ mol}^{-1}$$

□ **APPLICATION OF ENTHALPY OF COMBUSTION :**

Consider a reaction of disproportionation of C_6H_{14} into C_4H_8 and C_2H_6 . The following diagram show how enthalpy of combustions can be used to estimate enthalpy of reaction.



From the inspection of the above diagram, it is clear that

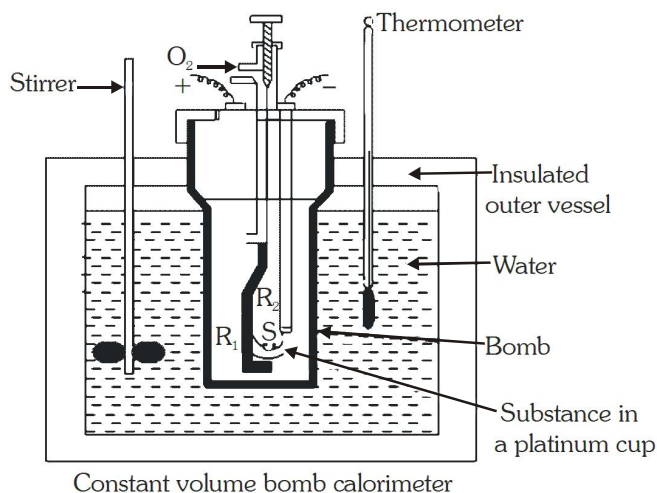
$$\Delta H_{\text{reaction}} = \Delta H_{\text{combustion}}(\text{C}_6\text{H}_{14}) - \Delta H_{\text{combustion}}(\text{C}_4\text{H}_8) - \Delta H_{\text{combustion}}(\text{C}_2\text{H}_6)$$

$$\text{Hence } \Delta H_{\text{reaction}} = \Delta H_{\text{combustion}}(\text{Reactant's}) - \Delta H_{\text{combustion}}(\text{Products})$$

❑ EXPERIMENTAL DETERMINATION OF ENTHALPY OF COMBUSTION :

(1) CONSTANT VOLUME BOMB CALORIMETER

The heat changes in chemical reactions are measured with the help of calorimeters. Depending upon the requirements of experiments different types of calorimeters are used. In general, reactions taking place at constant volume and involving gases are carried out in a closed container with rigid walls that can withstand high pressures such as a bomb calorimeter. The body of the bomb calorimeter is made of heavy steel. The steel vessel is coated inside with gold or platinum to avoid oxidation of steel during the chemical reactions. The vessel is fitted with a tight screw cap.



There are two electrodes R_1 and R_2 , which are connected to each other through a platinum wire S , which remains dipped in a platinum cup just below it. The substance under investigation is taken in platinum cup (few grams). The vessel is then filled with excess of oxygen at a pressure of about 15200 to 19000 mm of Hg and sealed. It is now dipped in an insulated water bath provided with a mechanical stirrer and a thermometer, sensitive enough to read upto 0.01 C (Beckmann's thermometer). The initial temperature of water is noted and the chemical reaction (combustion) is initiated by passing electric current through the platinum wire. The heat evolved during the chemical reactions raises the temperature of water, which is recorded from the thermometer. By knowing the heat capacity of the calorimeter and also the rise in temperature, the heat of chemical reaction or heat of combustion at constant volume can be calculated by using the expression.

$$\Delta U = -C \Delta T \quad (M/w)$$

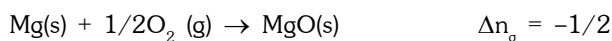
where, C = Heat capacity of calorimeter system ΔT = Rise in temperature M = Molecular mass of substance w = Mass of substance taken.

Ex. 0.138 gm sample of solid magnesium (molar mass = 24.3 g mol^{-1}) is burned in a constant volume bomb calorimeter that has a heat capacity of $1.77 \text{ kJ/}^\circ\text{C}$. The calorimeter contains 300 mL of water (density = 1 g/mL) and its temperature is raised by 1.126°C . Calculate enthalpy of combustion of magnesium at 289 K .

Sol. We have, $q_{\text{water}} = m_{\text{water}} C_{\text{water}} \Delta T$
 or $q_{\text{water}} = 300 \times 4.184 \times 1.126 = 1413 \text{ J} = 1.413 \text{ kJ}$
 $q_{\text{bomb}} = (m_{\text{bomb}} \cdot C_{\text{bomb}}) \Delta T$
 $= 1.77 \times 1.126 = 1.992 \text{ kJ}$
 $q_{\text{comb}} = -(q_{\text{water}} + q_{\text{bomb}})$
 $= -(1.413 + 1.992) = -3.405 \text{ kJ}$

$$\Delta U = \frac{q_{\text{comb}}}{\frac{m_{\text{subs}}}{M_{\text{subs}}}} = \frac{-3.405 \text{ kJ}}{\frac{0.138}{24.3}} = -599.6 \text{ kJ mol}^{-1}$$

The combustion reaction is



$$\begin{aligned} \Delta H &= \Delta U + (\Delta n_g) RT \\ &= -599.6 + (-1/2) (8.314 \times 10^{-3}) (298) \\ &= -599.6 - 1.24 = -600.84 \text{ kJ mol}^{-1} \end{aligned}$$

(2) CONSTANT PRESSURE BOMB CALORIMETER :

The only difference with the constant volume calorimeter is that the bomb is fitted with a piston which is maintained at constant pressure.

Thus the heat liberated during the chemical reaction q_p directly gives enthalpy of reaction by the expression

$$\Delta H = -C \Delta T \quad (M/w)$$

where, C = Heat capacity of calorimeter system, ΔT = Rise in temperature, M = Molecular mass of substance
w = Mass of substance taken.

□ MAXIMUM TEMPERATURE ATTAINED IN A BOMB CALORIMETER :

During explosion in bomb calorimeter, Heat released from the bond temporarily remain inside the bomb, and temperature of calorimeter system rises to a very high value. For small period the process can be considered as adiabatic process since the heat liberated still accumulates in container, and become available to product's formed during the reaction as well as remaining reactants.

$$\Delta_r U = \sum n_i C_{vi} (T_f - T_i) \quad \text{in a constant volume bomb calorimeter}$$

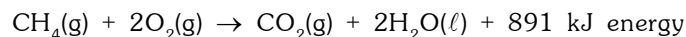
$$\Delta_r H = \sum n_i C_{pi} (T_f - T_i) \quad \text{in a constant pressure bomb calorimeter}$$

Here C_{pi} and C_{vi} are molar constant pressure and constant volume heat capacities of products and unreacted reactants and n_i are number of moles of respective substance.

□ CALORIFIC VALUE OF A FUEL

Nutritional value of foods, and the usefulness of fuels is determined by the heat (energy) they produce on heating. The energy released by the combustion of foods or fuels is usually compared in terms of their combustion energy per gram. It is known as calorific value. It is defined as the amount of heat produced in calories (or joules) when one gram of a substance (food or fuel) is completely burnt. The calorific value is usually expressed in kcal per gram or kilojoules per gram (1 kcal = 4.184 kJ).

For example when Methane burns



therefore calorific value of methane is $891/16 = 55.6 \text{ kJ/gm}$

The fuel are graded according to their calorific value. Greater the calorific value of a fuel, greater is the prize. Hydrogen has greatest calorific value.

□ BOND ENTHALPY

In chemical reactions the formation of a chemical bond is accompanied by the release of energy. Conversely energy has to be supplied for the breaking of a bond. Bond strengths are commonly described by their bond dissociation energy which is the energy required to break one mole of a bond of particular type. This is a definite quantity and is expressed in kJ mol^{-1} . For diatomic molecules the bond dissociation energy is same as bond energy, whereas in polyatomic molecules the bond energy is taken as the mean average of the various bond dissociation energies of the bonds of a given type.

Thus bond energy is average energy needed to break one mole of a particular bond in a gaseous molecule to convert it into gaseous atoms.

The thermochemical data is useful in determining the bond energies of different bonds.

For example the bond energy of C-H bond in methane can be calculated from its heat of formation. The heat for formation of methane from carbon and hydrogen gaseous atoms has been found to be $-1663 \text{ kJ mol}^{-1}$.



Methane has four C-H bonds and the energy required to break all the four C-H bonds is 1663 kJ. Therefore the average C-H bond energy is

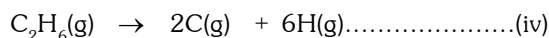
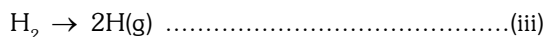
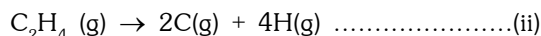
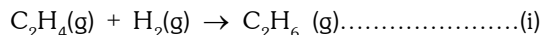
$$1663/4 = 413.25 \text{ kJ mol}^{-1}$$

APPLICATION OF BOND ENTHALPY IN ESTIMATION ENTHALPY OF REACTIONS

Bond energies can be utilized to quickly estimate enthalpy of reactions. The enthalpy obtained from bond energy data often match very well with actual enthalpy obtained from other methods like enthalpy of combustion, or formation enthalpy data.

$$\Delta H = \text{B.E.}(\text{reactant's}) - \text{B.E.}(\text{product's})$$

example : 1 - consider the reaction



Now (i) = (i) + (iii) - (iv)

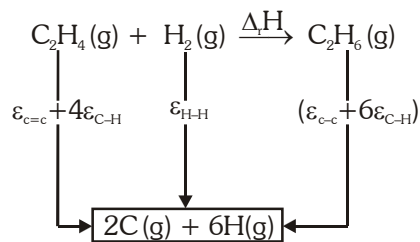
clearly from the equations the given formula is justified

Example :

2 - Figure in the following the enthalpies are indicated for.

$$\Delta_r H_{\text{Theoretical}} = \epsilon_{\text{C}=\text{C}} + 4\epsilon_{\text{C}-\text{H}} - 6\epsilon_{\text{C}-\text{C}} - \epsilon_{\text{C}-\text{H}}$$

$$\Rightarrow \Delta_r H_{\text{Theoretical}} = \text{B.E.}(\text{reactant's}) - \text{B.E.}(\text{product's})$$



RESONANCE ENERGY, CONCEPT AND CALCULATION USING BOND ENERGY :

When two or more double bond are in conjugation, there is possibility of delocalization of electron through conjugation. The Phenomenon is called resonance. Due to resonance, the molecule gain stability. The actual structure of molecule is average of many possible canonical structures possible for molecule. Resonance energy is difference in energy of most stable canonical structure and energy of actual molecule. When ever there is possibility of resonance in molecule, the molecule become more stable and bond breaking become difficult.

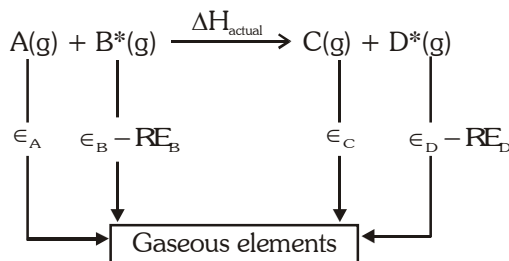
Resonance energy can be calculated using the formula

$$\Delta H(\text{Actual}) - \Delta H(\text{theoretical}) = \text{Resonance energy of products} - \text{Resonance energy of reactants}$$

The proof of above formula is given by following diagram.

Consider a reaction $\text{A}(\text{g}) + \text{B}^*(\text{g}) \rightarrow \text{C}(\text{g}) + \text{D}^*(\text{g})$ where (*) showing that molecules exhibit phenomena of resonance. Remember where ever resonance take place, bond breaking become difficult.

Actual energy required to break a bond is equal to $\epsilon_{\text{actual}} = \epsilon_{\text{theoretical}} - \text{resonance energy}$



$$\Delta H_{\text{actual}} = \epsilon_A + \epsilon_B - \text{R.E}_B - \{\epsilon_C + \epsilon_D - \text{R.E}_D\}$$

$$\Delta H_{\text{actual}} = (\epsilon_A + \epsilon_B - \epsilon_C - \epsilon_D) + \text{R.E}_D - \text{R.E}_B$$

$$\Rightarrow \Delta H_{\text{actual}} - \Delta H_{\text{Theoretical}} = \text{R.E}_D - \text{R.E}_B$$

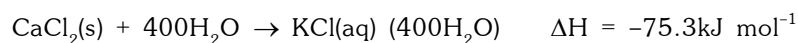
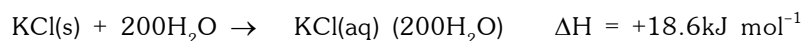
$$[\Delta H_{\text{actual}} - \Delta H_{\text{Theoretical}} = \text{R.E}_{\text{Products}} - \text{R.E}_{\text{reactants}}]$$

keep it in mind that resonance energy are expressed as -ive number because due to resonance, the molecule gain stability. Due to resonance in a molecule, bond breaking become difficult hence actual energy required to break a bond = theoretical bond energy - resonance energy.

□ **ENTHALPY OF SOLUTION :**

When a solute is dissolved in a solvent a solution is formed. During dissolution of a solute in any solvent, a certain amount of heat is either absorbed or evolved. Such heat changes under constant pressure conditions are known as the enthalpy of solution. 'The change in enthalpy when one mole of a solute is dissolved in a specified quantity of a solvent at a given temperature is called enthalpy of solution'. To avoid the amount of solvent, heat of solution is usually defined for an infinite dilute solution. Thus, heat of solution at infinite dilution is the heat change when one mole of a substance is dissolved in such a large quantity of solvent so that further dilution does not give any further heat change.

Here 'aq' represents aqueous meaning a large excess of water. For substances, which dissolve with the absorption of heat (endothermic), the enthalpy of solution is positive while for the substances which dissolve by liberating heat (exothermic), the enthalpy of solution is negative. For example, when KCl is dissolved in water, heat is absorbed. Thus, the enthalpy of solution of KCl is positive. For a 200 times dilution (water : KCl = 200 : 1), the enthalpy change during the process, So, the enthalpy of solution of KCl at a dilution of 200 is 18.6 kJ mol^{-1} . The dissolution of $\text{CaCl}_2(\text{s})$ in water is an exothermic process. So, the enthalpy of solution of calcium chloride (CaCl_2) is negative. At a dilution of 400, the enthalpy change for the reaction, So, the enthalpy of solution of $\text{CaCl}_2(\text{s})$ at a dilution of 400 is $-75.3 \text{ kJ mol}^{-1}$.



◆ **Important points to remember about enthalpy of solution :**

It is generally negative for anhydrous salts

It is generally positive for salts which do not have water of hydration like NaCl

It is generally positive for hydrated salts like $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$

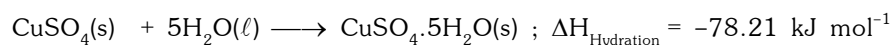
If lattice energy $|U|$ is more than $|\text{Hydration energy}|$ of gaseous ions, enthalpy of solution is positive.

If lattice energy $|U|$ is less than $|\text{Hydration energy}|$ of gaseous ions, enthalpy of solution is negative.

Note that even when $\Delta H_{\text{solution}}$ is positive, salts are freely soluble in water due to favourable Gibb's free energy change.

□ **ENTHALPY OF HYDRATION :**

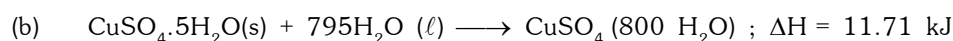
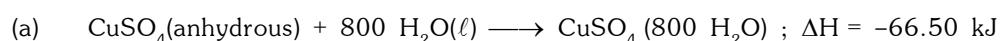
This is defined as the heat change (evolved or absorbed) when one mole of the anhydrous salt combines with the required number of moles of water to form the specific hydrated salt.



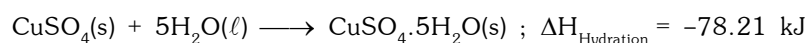
Hydration is generally exothermic change.

If integral heat of solution of the hydrated and anhydrous salt is known, then heat of hydration can be calculated.

For example :



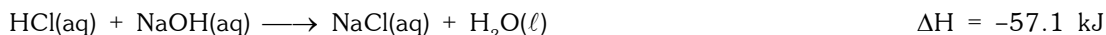
Thus, (a - b) gives



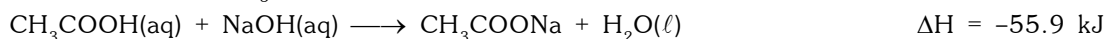
□ ENTHALPY OF NEUTRALISATION :

The reaction in which an acid and a base react to give a salt and water is called neutralization reaction. Neutralization reactions are exothermic in nature. The heat change when one gram equivalent of an acid is completely neutralised by a base or vice versa in dilute solution, is called heat of neutralization. Examples of heat of neutralization are :

Neutralization of HCl with NaOH



Neutralization of CH_3COOH with NaOH

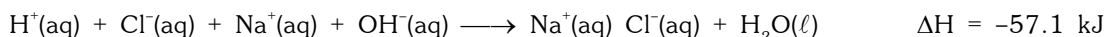


It is important to note that the term gram equivalent is used in the definition of heat of neutralization. This is because neutralization involves 1 mole of H^+ ions and 1 mole of OH^- ions to form 1mole of water and 57.1 kJ of heat is liberated.



Now, one gram equivalent of various acids on complete dissociation liberates one mole of H^+ ions. But one mole of the acid may produce more than one mole of H^+ ions in solution depending upon its basicity; for example 1mol of H_2SO_4 gives 2 mol of H^+ ions and 1mol of H_3PO_4 gives 3 mol of H^+ ions on complete dissociation. But 1gram equivalent of both (H_2SO_4 or H_3PO_4) produces only 1 mol of H^+ ions. Thus, it is more appropriate to use the term gram equivalent in the definition of enthalpy of neutralization.

The average enthalpy of neutralization of any strong acid by a strong base is found to be -57.7 . This is because strong acids and strong bases are completely ionized in aqueous solutions. The aqueous solution of one gram equivalent of all strong acids contains the same number of H^+ ions. Similarly, aqueous solution of one gram equivalent of all strong bases also contains same number of OH^- . The neutralization reactions between strong acids and strong bases in aqueous solutions involve simply the combination of H^+ ions (from an acid) and OH^- ions (from a base) to form unionized water molecules. For example, the reaction between hydrochloric acid and sodium hydroxide. The neutralization can be represented as :



Cancelling common ions :



□ EXPERIMENTAL DETERMINATION OF ENTHALPY OF NEUTRALIZATION :

In this experiment, an aqueous hydrochloric acid solution will be added to an aqueous sodium hydroxide solution. The neutralization reaction will occur until either H^+ or OH^- is entirely consumed. The reactant which is consumed completely is called the limiting reactant.

The molar enthalpy of neutralization is defined as

$$\Delta H_{\text{neut}} = q_{\text{neut}}/n$$

Where q_{neut} is the heat of neutralization, measured calorimetrically, and n is the moles of the limiting reactant.

Approach \therefore Add a known volume of 3.00 M aqueous HCl to a known volume of 1.00 M aqueous NaOH.

- ◆ Observe the temperature of the system before and after the neutralization reaction occurs.
- ◆ Calculate the change in temperature (ΔT) for the system.
- ◆ Use the temperature change and heat capacities (C) for the calorimeter and aqueous solution to calculate the heat of neutralization.
- ◆ Divide the heat of neutralization by the moles of the limiting reactant to determine the molar enthalpy of neutralization.

$$\Delta H_{\text{neut}} = -\frac{C\Delta T}{n}$$

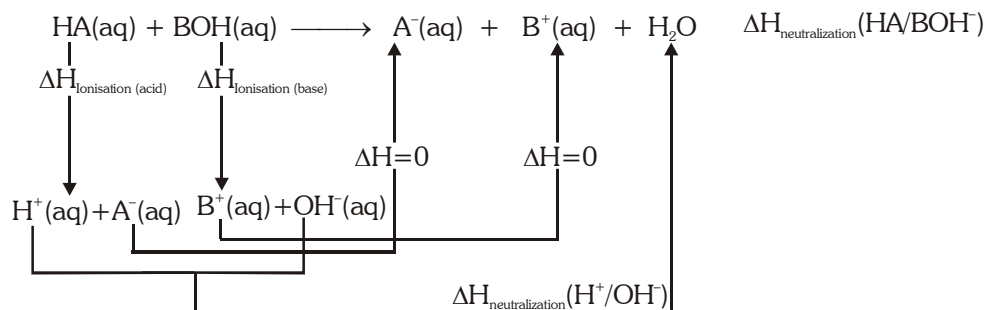
THE ENTHALPY OF NEUTRALIZATION OF WEAK ACID AND WEAK BASE :

Consider the neutralisation of weak acid HA with weak base BOH.

Where, $\Delta H_{\text{ionisation}}(\text{HA})$ = Enthalpy of ionization of acid HA

$\Delta H_{\text{ionisation}}(\text{BOH})$ = Enthalpy of ionization of base BOH

$\Delta H_{\text{neutralisation}}(\text{H}^+/\text{OH}^-)$ = Enthalpy change for the reaction of H^+ and OH^- to form water.

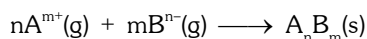
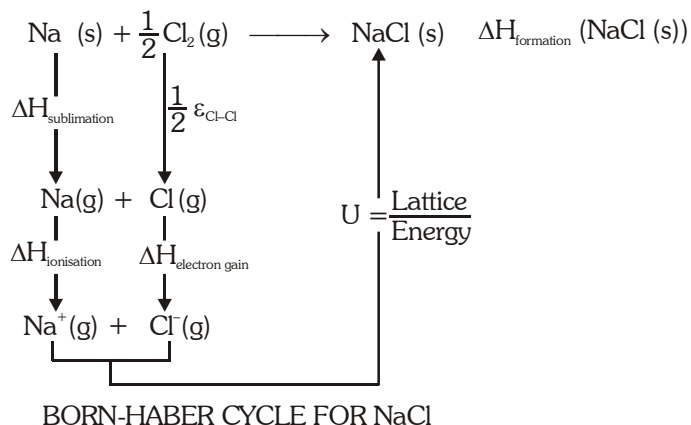


Thus for weak acid :

$$\Delta H_{\text{neut}}(\text{BOH/HA}) = \Delta H_{\text{ion.}}(\text{HA}) + \Delta H_{\text{ion}}(\text{BOH}) + \Delta H_{\text{neut}}(\text{H}^+ \text{OH}^-)$$

BORN - HABER CYCLE FOR NaCl :

This cycle is based on thermochemical changes taking place in the formation of a lattice. This cycle can be used to determine lattice energy which cannot be directly measured. It is defined as that energy released when one mole of the ionic compound (lattice) is formed from its isolated ions in the gaseous state under standard condition.



$$\Delta H = -U \text{ (lattice energy)}$$

Formation of NaCl(s) lattice involves thus,

$$S + I + \frac{\epsilon_{\text{Cl-Cl}}}{2} - E - U = q$$

hence, U can be calculated.

here, S = enthalpy of sublimation of Na(s) = $\Delta H_{\text{sublimation}}$

I = ionisation of energy of Na(g) = $\Delta H_{\text{ionization}}$

ϵ = bond energy of Cl_2

U = lattice energy

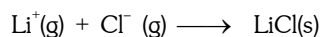
q = enthalpy of formation of NaCl(s) = $\Delta H_{\text{formation}}$

If lattice is $\text{MgX}_2(\text{s})$ then

$$S + (I_1 + I_2) + \varepsilon - 2E - U = q$$

where, $(I_1 + I_2)$ = total ionisation energy to form $\text{Mg}^{2+}(\text{g})$.

Ex. Calculate lattice energy for the change,



Given that

$\Delta H_{\text{subl.}}$ of Li = $160.67 \text{ kJ mol}^{-1}$,

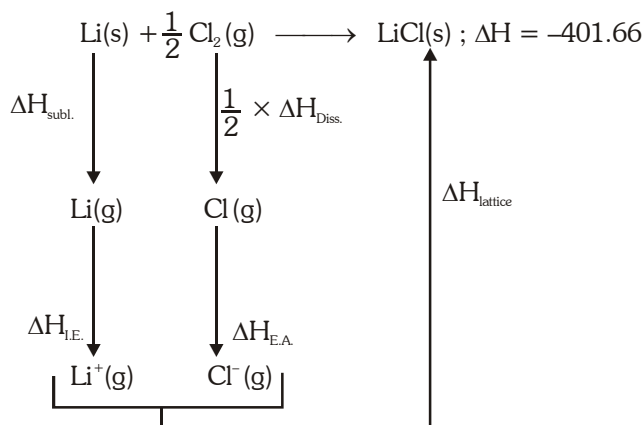
$\Delta H_{\text{Dissociation}}$ of Cl_2 = $244.34 \text{ kJ mol}^{-1}$,

$\Delta H_{\text{ionisation}}$ of Li(g) = $520.07 \text{ kJ mol}^{-1}$,

$\Delta H_{\text{E.A.}}$ of Cl(g) = $-365.26 \text{ kJ mol}^{-1}$,

ΔH_f of LiCl(s) = $-401.66 \text{ kJ mol}^{-1}$.

Sol. Considering the different changes that occur in the formation of solid lithium chloride based on the data given the lattice energy of the above can be constituted as:



$$\Delta H_f = \Delta H_{\text{subl.}} + \Delta H_{\text{I.E.}} + \frac{1}{2} \Delta H_{\text{Diss.}} + \Delta H_{\text{E.A.}} + \Delta H_{\text{lattice}}$$

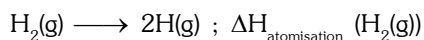
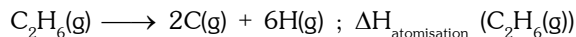
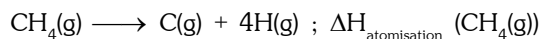
or
$$\Delta H_{\text{lattice}} = \Delta H_f - \Delta H_{\text{subl.}} - \Delta H_{\text{I.E.}} - \frac{1}{2} \Delta H_{\text{Diss.}} - \Delta H_{\text{E.A.}}$$

$$= -839.31 \text{ kJ mol}^{-1}$$

□ **ENTHALPY OF ATOMISATION :**

Enthalpy change to convert 1 mole of molecule into gaseous atoms is called enthalpy of atomisation :

Example :



Ex. The bond dissociation enthalpies of $\text{H}_2(\text{g})$ and $\text{N}_2(\text{g})$ are $+435.95 \text{ kJ mol}^{-1}$ and $+941.8 \text{ kJ mol}^{-1}$ and enthalpy of formation of $\text{NH}_3(\text{g})$ is $-46.024 \text{ kJ mol}^{-1}$.

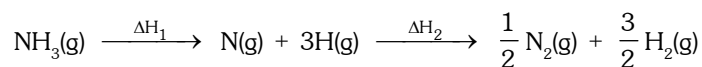
(a) What is the enthalpy of atomization of $\text{NH}_3(\text{g})$?

(b) What is the bond enthalpy of N – H bond ?

Sol. (a) The desired equation is



we can consider the reaction as



$$\therefore \Delta H_1 + \Delta H_2 = +46.024 \text{ kJ mol}^{-1}$$

$$\text{Now } \Delta H_2 = -\frac{3}{2} (435.95) - \frac{1}{2} (941.80)$$

$$\therefore \Delta H_1 = 46.024 + \frac{3}{2} (435.925) + \frac{1}{2} (941.8)$$

$$\text{or } \Delta H_1 = 46.024 + 653.925 + 470.9$$

$$\text{or } \Delta H_1 = 1170.849 \text{ kJ mol}^{-1}$$

$$\text{or } \Delta H_1 = 1.170849 \text{ MJ mol}^{-1}$$

(b) The average bond enthalpy of N – H bond

$$= \frac{1170.849}{3} = 390.283 \text{ kJ mol}^{-1}$$

□ ENTHALPY OF ELECTRON GAIN :

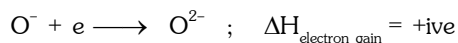
Enthalpy change when 1 mole electrons are added to gaseous species is called electron gain enthalpy.

Example :



$\Delta H_{\text{electron gain}}$ is usually negative for non metals.

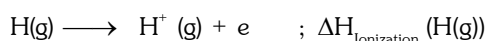
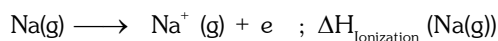
However electron gain enthalpy of negatively charged gaseous species is positive.



□ IONISATION ENTHALPY :

Enthalpy change when one mole of gaseous atom is converted into gaseous ion by removing one mole electron from ground state is called ionisation enthalpy.

Example :



□ Different types of standard enthalpy changes (Based on IUPAC recommendations).

	Transition Enthalpy	Example	Definition	Symbol
1.	Enthalpy of formation	$\text{K(s)} + \frac{1}{2} \text{Cl}_2(\text{g}) \longrightarrow \text{KCl(s)}$	One mole of the compound formed from its elements	$\Delta_f H$ or ΔH_f
2.	Enthalpy of combustion	$\text{C}_2\text{H}_6(\text{g}) + \text{O}_2(\text{g}) \longrightarrow \text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\ell)$	One mole of the compound (fuel) is burnt completely in oxygen	$\Delta_c H$ or ΔH_c
3.	Enthalpy of fusion	$\text{H}_2\text{O(s)} \longrightarrow \text{H}_2\text{O}(\ell)$	One mole of the liquid is formed from the solid without a change in temperature	$\Delta_{fus} H$ or ΔH_{fus}
4.	Enthalpy of vaporisation	$\text{H}_2\text{O}(\ell) \longrightarrow \text{H}_2\text{O(g)}$	One mole of the vapour formed from the liquid without a change in temperature	$\Delta_{vap} H$ or ΔH_{vap}
5.	Enthalpy of sublimation	$\text{I}_2(\text{s}) \longrightarrow \text{I}_2(\text{g})$	One mole of the vapour formed from the solid	$\Delta_{sub} H$ or ΔH_{sub}
6.	Enthalpy of atomisation	$\text{H}_2(\text{g}) \longrightarrow 2\text{H(g)}$	One mole of the substance broken into isolated atoms in the gas phase	$\Delta_{at} H$ or ΔH_{at}
7.	Enthalpy of reaction	$\text{A} \longrightarrow \text{B}$	Enthalpy change taking place in a reaction	$\Delta_r H$ or ΔH_r
8.	Enthalpy of neutralisation	$\text{HCl(aq)} + \text{NaOH(aq)} \longrightarrow \text{H}_2\text{O} + \text{NaCl(aq)}$	One mole of water formed by the neutralisation of an acid by a base	$\Delta_n H$ or ΔH_n
9.	Enthalpy of ionisation	$\text{Na(g)} \longrightarrow \text{Na}^+(\text{g}) + \text{e}^-$	One mole atoms ionised - all species in gaseous phase	$\Delta_{ion} H$ or ΔH_{ion}
10	Electron-gain enthalpy	$\text{X(g)} + \text{e}^- \longrightarrow \text{X}^-(\text{g})$	One mole of anions being formed all species in the gaseous phase	$\Delta_{eg} H$ or ΔH_{eg}
11	Lattice enthalpy	$\text{NaCl(s)} \longrightarrow \text{Na}^+(\text{g}) + \text{Cl}^-(\text{g})$	One mole of a crystal completely separated into isolated particles in the gaseous phase	$\Delta_L H$ or ΔH_L or U
12	Bond dissociation enthalpy (bond A - B)	$\text{HCl(g)} \longrightarrow \text{H(g)} + \text{Cl(g)}$	One mole of bonds broken - all species in the gaseous phase	$\Delta H_{A-B} = \epsilon_{A-B}$
13	Enthalpy of solution	$\text{NaCl(s)} + \text{H}_2\text{O(excess)} \longrightarrow \text{NaCl(aq)}$	One mole of the solute dissolved in excess of solvent so that further dilution produces no enthalpy change	$\Delta_{sol} H$ or ΔH_{sol}
14	Enthalpy of hydration	$\text{X}^+(\text{g}) \longrightarrow \text{X}^+(\text{aq})$	One mole of the ion in gaseous phase is hydrated	$\Delta_{hyd} H$ or ΔH_{hyd}
15	Enthalpy of mixing	pure substances \rightarrow mixture	One mole each of the two or more substances is mixed	$\Delta_{mix} H$ or ΔH_{mix}

MEMORY TIPS

- ◆ Heat change expressed at constant pressure = ΔH

Heat change expressed at constant volume = ΔE

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

$$\Delta H = \Delta E + \Delta nRT$$

Δ = No. of moles of product – No. of moles of reactants ; as represented by stoichiometry of change.

(count only gaseous phase molecule for Δn)

- ◆ **Kirchoff's Equation** : Variation of ΔH and ΔE with temperature

$$\Delta H_2 - \Delta H_1 = \Delta C_p (T_2 - T_1)$$

$$\Delta E_2 - \Delta E_1 = \Delta C_v (T_2 - T_1)$$

$$\Delta C_p = \sum C_{p \text{ Products}} - \sum C_{p \text{ Reactants}}$$

$$\Delta C_v = \sum C_{v \text{ Products}} - \sum C_{v \text{ Reactants}}$$

ΔC_p and ΔC_v are changes in molar heat capacities at constant and volume respectively during the change.

- ◆ **Standard Heat Enthalpy (H)**

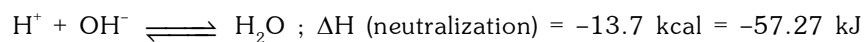
$$H_{(\text{Compound})} = \Delta H_{(\text{Formation of compound})}$$

- ◆ **Heat of solution for electrolytes (ΔH_s)**

$$\Delta H_s = \Delta H_i + \Delta H_h$$

ΔH_i and ΔH_h are heat of ionisation and heat of hydration respectively

- ◆ **Heat of neutralization**

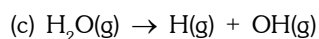
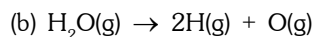
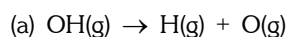


SOLVED PROBLEMS (OBJECTIVE)

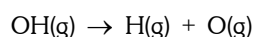
Ex.1 From the following data at 25°C

Reaction	$\Delta_r H^\circ / \text{KJ mol}^{-1}$
$1/2 \text{H}_2(\text{g}) + 1/2 \text{O}_2(\text{g}) \rightarrow \text{OH}(\text{g})$	42.09
$\text{H}_2(\text{g}) + 1/2 \text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{g})$	-241.84
$\text{H}_2(\text{g}) \rightarrow 2\text{H}(\text{g})$	435.88
$\text{O}_2(\text{g}) \rightarrow 2\text{O}(\text{g})$	495.04

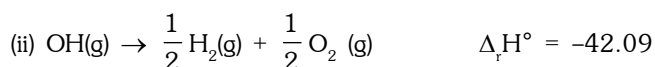
Calculate $\Delta_r H^\circ$ for the following reactions



Sol. (a) The desired equation is

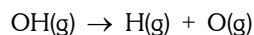


We are to develop the desired equation by using the four given equations



By adding equation (i), (ii) & (iii),

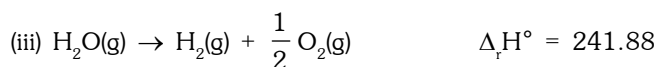
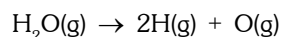
we get



$$\Delta_r H^\circ \rightarrow \frac{435.88}{2} - 42.09 + \frac{495.04}{2}$$

or $\Delta_r H^\circ = 423.37 \text{ kJ/mol}$. **Ans.**

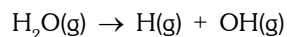
(b) The desired equation is



The net equation is,



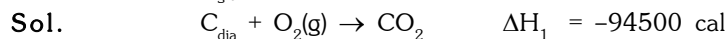
(c) The desired equation is



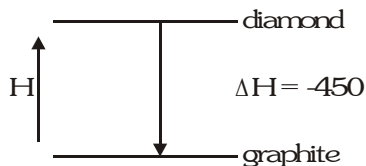
By adding eqⁿ (i) and (ii)



Ex.2 From the given data prove that graphite is more stable allotrope of carbon than diamond from the given data



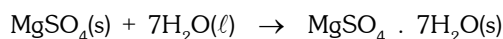
on subtraction



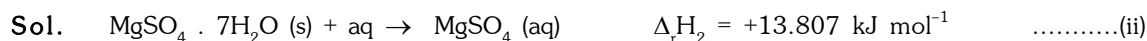
since heat content of diamond > graphite

Therefore diamond is less stable than graphite.

Ex.3 At 25° C, 1 mole $MgSO_4$ was dissolved in water. The heat evolved was found to be 91.211 kJ. One mole of $MgSO_4 \cdot 7H_2O$ on dissolution gives a solution of the same composition accompanied by an absorption of 13.807 kJ. Find the enthalpy of hydration, i.e., ΔH for the reaction



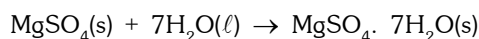
Given that



Equation (i) can be written as follows :



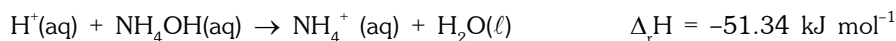
eqn (iii) – (ii) will give



$$\Delta_r H = -91.211 - 13.807 = -105.018 \text{ kJ mol}^{-1}.$$

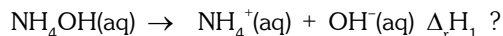
Ex.4 Enthalpy of neutralization of HCl by NaOH is $-57.32 \text{ kJ mol}^{-1}$ and by NH_4OH is $-51.34 \text{ kJ mol}^{-1}$. Calculate the enthalpy of dissociation of NH_4OH .

Sol. Given that



we may consider neutralization in two steps :

(i) ionization



(ii) neutralization

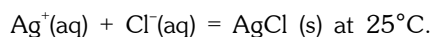


$$\text{Thus, } \Delta_r H = \Delta_r H_1 + \Delta_r H_2$$

Therefore,

$$\begin{aligned} \Delta_r H_1 &= \Delta_r H - \Delta_r H_2 \\ &= -51.34 + 57.32 = 5.98 \text{ kJ/mol}^{-1} \end{aligned}$$

Ex.5 Calculate $\Delta_r H^\circ$ for the reaction



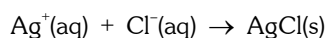
Given :

$$\Delta_f H^\circ (Ag^+, aq) = 105.58 \text{ kJ mol}^{-1},$$

$$\Delta_f H^\circ (Cl^-, aq) = -167.16 \text{ kJ mol}^{-1} \text{ and } \Delta_f H^\circ$$

$$(AgCl, s) = -127.07 \text{ kJ mol}^{-1}$$

Sol. For the reaction



We have

$$\begin{aligned}\Delta_f H^\circ &= (\text{AgCl, s}) - \Delta_f H^\circ (\text{Ag}^+, \text{aq}) - \Delta_f H^\circ (\text{Cl}^-, \text{aq}) \\ &= [-127.07 - 105.58 - (-167.16)] \\ &= -65.49 \text{ kJ mol}^{-1}\end{aligned}$$

Ex.6 The enthalpy of formation of ethane, ethylene and benzene from the gaseous atoms are -2839.2 , -2275.2 and $-5536 \text{ kJ mol}^{-1}$ respectively. Calculate the resonance energy of benzene, compared with Kekule structure. The bond enthalpy of C – H bond is given as equal to $410.87 \text{ kJ mol}^{-1}$.

Sol. Bond enthalpy of C – C bond

$$\begin{aligned}&= \text{Enthalpy required to break } \text{C}_2\text{H}_6 \text{ into gaseous atoms} - 6 \text{ bond enthalpy of C – H bond} \\ &= 2839.2 \text{ kJ mol}^{-1} - 6 \times 410.87 \text{ kJ mol}^{-1} \\ &= 373.98 \text{ kJ mol}^{-1}\end{aligned}$$

Bond enthalpy of C = C bond = Enthalpy required to break C_2H_4 into gaseous atoms $-4 \times$ bond enthalpy of C – H bond

$$\begin{aligned}&= 2275.2 \text{ kJ mol}^{-1} - 4 \times 410.87 \text{ kJ mol}^{-1} \\ &= 631.72 \text{ kJ mol}^{-1}\end{aligned}$$

For the formation of benzene having Kekule structure, we have to form 3 C – C bonds, 3 C = C bonds and 6 C – H bonds for which enthalpy released is

$$\begin{aligned}&[3(-373.98) + 3(-631.72) + 6(-410.87)] \\ &= -5482.32 \text{ kJ mol}^{-1}\end{aligned}$$

But the given value of $\Delta_f H$ is

$$\Delta_f H (\text{actual}) = -5536 \text{ kJ mol}^{-1}$$

Hence resonance energy compared to Kekule structure

$$\begin{aligned}&= \Delta_f H (\text{actual}) - \Delta_f H (\text{Kekule structure}) \\ &= (-5536 + 5482.32) \\ &= -53.68 \text{ kJ mol}^{-1}\end{aligned}$$

Ex.7 Using bond enthalpy data given below, estimate enthalpy of formation of acetic acid.

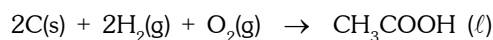
Bond	Bond enthalpies
C – H	$413.38 \text{ kJ mol}^{-1}$
C – C	$347.69 \text{ kJ mol}^{-1}$
C = O	$728.02 \text{ kJ mol}^{-1}$
C – O	$351.46 \text{ kJ mol}^{-1}$
O – H	$462.75 \text{ kJ mol}^{-1}$

Enthalpy of atomization

C	$718.39 \text{ kJ atom}^{-1}$
H	$217.94 \text{ kJ atom}^{-1}$
O	$247.52 \text{ kJ atom}^{-1}$

The observed $\Delta_f H^\circ$ for acetic acid is $-438.15 \text{ kJ mol}^{-1}$. Compute the resonance energy of acetic acid.

Sol. The desired reaction for the formation of CH_3COOH is



$$\begin{aligned}\Delta_f H^\circ &= -3 \times 413.38 - 1 \times 347.69 - 728.02 - 351.46 - 462.75 + 2 \times 718.39 + 4 \times 217.94 + 2 \times 247.52 \\ &= -326.48 \text{ kJ mol}^{-1}\end{aligned}$$

$$\text{Resonance energy} = \Delta H^\circ_{\text{obs}} - \Delta H^\circ$$

$$= -438.15 + 326.48$$

$$= -111.67 \text{ kJ mol}^{-1}$$

Ex.8 The specific heats of iodine vapour and solid are 0.031 and 0.055 cal/g respectively. If heats of sublimation of iodine is 24 cal/g at 200°C, what is its value at 250°C ?

Sol. Given



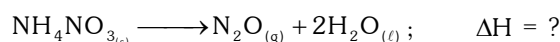
$$\begin{aligned} \Delta C_{p(\text{cal/g})} &= C_p \text{ of product} - C_p \text{ of reactant} \\ &= 0.031 - 0.055 \\ &= -0.024 \text{ cal/g} \end{aligned}$$

$$\begin{aligned} \text{Now } \Delta H_2 - \Delta H_1 &= \Delta C_p (T_2 - T_1) \\ \Delta H_2 - 24 &= -0.024 (523 - 473) \\ \therefore \Delta H_2 &= 24 - 1.2 = 22.8 \text{ cal/g.} \end{aligned}$$

Ex.9 The molar heat of formation of $\text{NH}_4\text{NO}_{3(s)}$ is -367.54 kJ and those of $\text{N}_2\text{O}_{(g)}$, $\text{H}_2\text{O}_{(l)}$ are 81.46 and -285.8 kJ respectively at 25°C and 1 atmosphere pressure. Calculate ΔH and ΔE of the reaction

$$\text{NH}_4\text{NO}_{3(s)} \longrightarrow \text{N}_2\text{O}_{(g)} + 2\text{H}_2\text{O}_{(l)}$$

Sol. We have to find ΔH for



$$\begin{aligned} \Delta H_{\text{reaction}} &= \Delta H_{\text{Products}} - \Delta H_{\text{Reactants}} \\ &= \Delta H_{\text{N}_2\text{O}} + \Delta H_{\text{H}_2\text{O}} \times 2 - \Delta H_{\text{NH}_4\text{NO}_3} \end{aligned}$$

$$\text{Given, } \Delta H_{\text{N}_2\text{O}} = +81.46 \text{ kJ, } \Delta H_{\text{H}_2\text{O}} = -285.8 \text{ kJ, } \Delta H_{\text{NH}_4\text{NO}_3} = -367.54 \text{ kJ}$$

$$\begin{aligned} \therefore \Delta H_{\text{reaction}} &= +81.46 + 2(-285.8) - (-367.54) \\ \Delta H &= -122.6 \text{ kJ} \end{aligned}$$

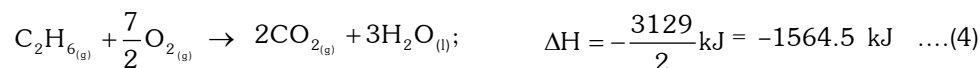
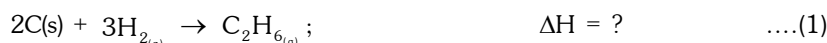
$$\text{Further } \Delta H = \Delta E + \Delta nRT \quad (\Delta n = 1 - 0 = 1, R = 8.314 \text{ J, } T = 298 \text{ K})$$

$$\therefore -122.6 \times 10^3 = \Delta E + 1 \times 8.314 \times 298$$

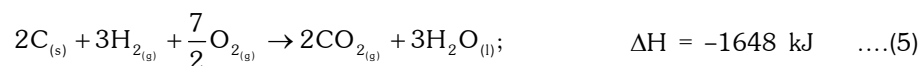
$$\begin{aligned} \therefore \Delta E &= -125077 \text{ joule} \\ &= -125.077 \text{ kJ} \end{aligned}$$

Ex.10 When 2 mole of $\text{C}_2\text{H}_{6(g)}$ are completely burnt, 3129 kJ of heat is liberated. Calculate the heat of formation of $\text{C}_2\text{H}_{6(g)}$. ΔH_f for $\text{CO}_{2(g)}$ and $\text{H}_2\text{O}_{(l)}$ are -395 and -286 kJ mol^{-1} respectively.

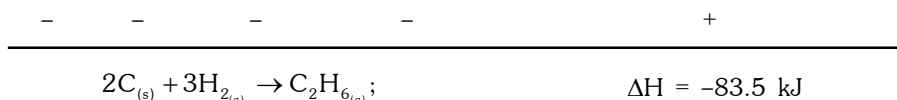
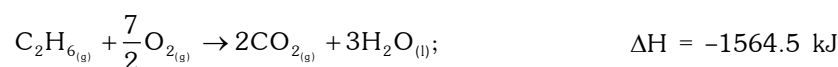
Sol. We have to find



Multiplying Eq. (2) by 2 and Eq. (3) by 3, then adding



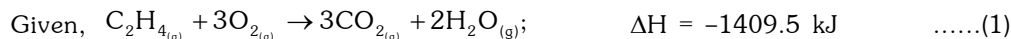
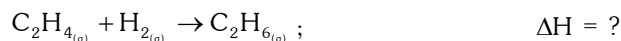
Subtracting Eq. (4) from Eq. (5)



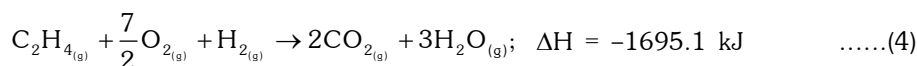
$$\therefore \Delta H_f \text{ of } \text{C}_2\text{H}_6 = -83.5 \text{ kJ}$$

Ex.11 The heats of combustion of $C_2H_{4(g)}$, $C_2H_{6(g)}$ and $H_{2(g)}$ are -1409.5 , -1558.3 and -285.6 kJ respectively. Calculate heat of hydrogenation of ethylene.

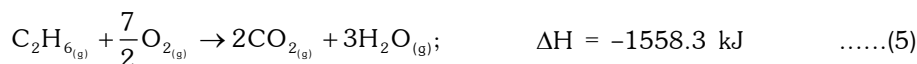
Sol. We have to find



Adding Eqs. (1) and (3)



Subtracting Eq. (2) from (4)



- - - - +



\therefore Heat of hydrogenation of $C_2H_4 = 136.8 \text{ kJ}$

Ex.12 ΔH for combustion of ethane and ethyne are -341.1 and -310.0 kcal respectively. Which is better gas welder and why ?

Sol. A better gas welder is one which posses high calorific value, i.e., heat produced by 1 g of fuel.

$\Delta H_{\text{combustion}}$ for $C_2H_6 = -341.1 \text{ kcal}$

$$\therefore \text{Calorific value} = \frac{-341.1}{30} \text{ kcal/g} = -11.37 \text{ kcal/g}$$

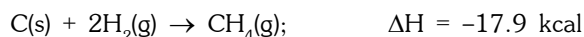
$\Delta H_{\text{combustion}}$ for $C_2H_2 = -310.0 \text{ kcal}$

$$\therefore \text{Calorific value} = \frac{-310.0}{26} \text{ kcal/g} = -11.92 \text{ kcal/g}$$

$\therefore C_2H_2$ is better gas welder.

Ex.13 The heat of formation of methane is -17.9 kcal. If the heats of atomisation of carbon and hydrogen are 170.9 and 52.1 kcal per mole, calculate the C-H bond energy in methane.

Sol. Given that



Energy change in reactants :

Heat of atomisation of 1 mole of C = 170.9 kcal

Heat of atomisation of 4 moles of H = $4 \times 52.1 \text{ kcal}$

Energy change in product :

Heat of formation of 4 moles of C-H bonds = $4 \times x \text{ kcal}$.

(where x is the energy of formation of C-H bonds in kcal/mole).

Since the algebraic sum of all the heat changes is equal to the heat of formation of the above given equation, we have

$$170.9 + 4 \times 52.1 + 4x = -17.9; \quad x = -99.3 \text{ kcal}$$

Thus the bond energy = $+99.3 \text{ kcal/mole}$.

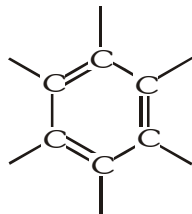
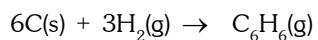
Ex. 14 Calculate the heat of formation of benzene from the following data, assuming no resonance. Bond energies :

$$\text{C}-\text{C} = 83 \text{ kcal}, \quad \text{C}=\text{C} = 140 \text{ kcal}, \quad \text{C}-\text{H} = 99 \text{ kcal}$$

Heat of atomisation of C = 170.9 kcal

Heat of atomisation of H = 52.1 kcal

Sol. We have to calculate ΔH for the reaction



$\Delta H = ?$

For reactants :

Heat of atomisation of 6 moles of C = 6 \times 170.9 kcal

Heat of atomisation of 6 moles of H = 6 \times 52.1 kcal

For products :

Heat of formation of 6 moles of C - H bonds = -6 \times 99

Heat of formation of 3 moles of C - C bonds = -3 \times 83

Heat of formation of 3 moles of C = C bonds = -3 \times 140

On adding, we get heat of formation of C_6H_6 , i.e.,

$$\Delta H = 6 \times 170.9 + 6 \times 52.1 - 6 \times 99 - 3 \times 83 - 3 \times 140 = 75.0 \text{ kcal.}$$
