Enthalpy changes of chemical Reactions: Reaction Enthalpy:

The enthalpy of reaction is defined as —
The enthalpy change in a reaction when the number of moles of reactants react completely to give the product as given by the balanced chemical equation.

For example, the enthalpy change for the reaction of one mole of methane with two moles of ~~water~~ is oxygen to form one mole of CO_2 and two moles of water is -890.3 KJ.

Thus,

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$$

$$\Delta_r H = -890.3 \text{ KJ.}$$

Enthalpy changes of Reactions: In a reaction

$$\text{Reactants} \longrightarrow \text{Products}$$

Enthalpy change of a reaction $\Delta_r H$ is given by

$$\Delta_r H = \left[\text{Sum of enthalpies of Products}\right] - \left[\text{Sum of enthalpies of reactants}\right]$$

$$\Delta_r H = \Sigma \alpha_i H_{\text{Product}} - \Sigma \beta_i H_{\text{reactants}}$$

Ques. Red Phosphorus [reacts] with liquid bromine as:

$$2P(s) + 3Br_2(l) \longrightarrow 2PBr_3(g) \quad \Delta_r H^\circ = -243 \text{ kJ}$$

Calculate the enthalpy change when 10.32 g of Phosphorus reacts with an excess of Bromine.

Solution: $2P_{(s)} + 3Br_{2(l)} \longrightarrow 2PBr_{3(g)}$

$\Delta_r H^\circ = -243\ KJ\,mol^{-1}$

Molar mass of phosphorous $= 31\ g\,mol^{-1}$

Mass of $P = 10.32\ g$

Moles of $P = \frac{10.32}{31} = 0.333\ mol$

Enthalpy change for 2 moles of $P = -243\ KJ$

1 ——— $= -\frac{243}{2}\ KJ$

0.333 ——— $= -\frac{243}{2} \times 0.333$

$= -40.46\ KJ.$

Enthalpy of formation: The enthalpy change when 1 mole of a substance is formed from its elements in their most stable forms is called enthalpy of formation. for example. the formatin of CO_2 and [redacted] may be expressed as:

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

It may be noted that these equations should always be written as for one mole of the substance to be formed.

Que. Calculate the enthalpy change in the reaction:

$$4NH_3(g) + 3O_2(g) \longrightarrow 2N_2(g) + 6H_2O(l)$$

at 298K given that the enthalpy of formation at 298K for $NH_3(g)$ and $H_2O(l)$ are -46.0 and $-286.0\ KJmol^{-1}$ respectively.

Solution.: $\Delta_r H^\circ$ for the reaction,

$$4NH_3(g) + 3O_2(g) \longrightarrow 6H_2O(l) + 2N_2(g)$$

$$\Delta_r H^\circ = \Delta_f H^\circ(\text{product}) - \Delta_f H^\circ(\text{reactants})$$

$$\Delta_r H^\circ = \{6\Delta_f H^\circ[H_2O_{(l)}] + 2\Delta_f H^\circ[N_2(g)]\} - \{4\Delta_f H^\circ[NH_3(g)] + 3\Delta_f H^\circ[O_2]\}$$

$$\Delta_r H^\circ = \{6 \times (-286) + 0\} - \{4 \times (-46) + 0\}$$

$$\Delta_r H^\circ = -1716 + 184$$

$$\boxed{\Delta_r H^\circ = -1532\ KJ}$$

Enthalpy of Combustion: The enthalpy change when 1 mole of a substance is completely burnt in excess of oxygen or air is called enthalpy of combustion. for example.

① Combustion of Carbon

$$C_{(s)} + O_2(g) \longrightarrow CO_2(g)$$

$$\Delta H = \Delta_c H = -393.5\ KJ$$

② Combustion of methane.:

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$$

$$\Delta H = \Delta_c H = -890.3\ KJ$$