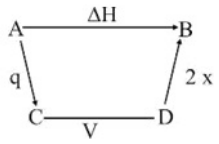
Daily Tutorial Sheet-3	Level-1

31.(C) According to Hess's law, the total heat changes occurring during a chemical reaction are independent of path.



$\Delta H = q + V + 2x$

32.(D) $CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$ $\Delta n_g = 1 - 3 = -2$

We know that, $\Delta H = \Delta E + \Delta n_g RT$

$\therefore$ $\Delta H = (-885389) - (-2) \times 8.314 \times 298 = -885389 - 4955.1440 = -890344.4\,J\,mol^{-1}$

33.(C) Relation between ΔH (enthalpy change) and ΔE (internal energy change) is: $\Delta H = \Delta E + \Delta_{n_g} RT$

Where, Δ_{n_g} = moles of gaseous products-moles of gaseous reactants $= 2 - 3 = -1$

$\Rightarrow$ $-1366.5 = \Delta E - 1 \times 8.314 \times 10^{-3} \times 300$

$\Delta E = -1364.0$ kJ mol^{-1}

34.(A) According to Hess's law total heat changes during a chemical reaction are independent of path of reaction.

Given, $I_2(s) \rightarrow I_2(g)$, $\Delta H_1 = 57.3$ kJ / mol ...(i)

$I_2(s) \rightarrow I_2(\ell)$, $\Delta H_2 = +15.5$ kJ / mol ...(ii)

Required equation $I_2(l) \rightarrow I_2(g)$, $\Delta H = ?$ subtract Eq. (ii) from Eq. (i)

$\therefore$ $I_2(l) \longrightarrow I_2(g)$, $\Delta H = 57.3 + (-15.5) = +41.8$ kJ / mol

35.(B) Given, $\frac{1}{2}N_2(g) + \frac{3}{2}H_2(g) \rightarrow NH_3(g)$

$\Delta H_f^\circ = -46.0$ kJ mol^{-1}

$2H(g) \rightarrow H_2(g)$ $\Delta H_f^\circ = -436$ kJ mol^{-1}

$2N(g) \rightarrow N_2(g)$ $\Delta H_f^\circ = -712$ kJ mol^{-1}

Assuming X is the bond energy of N – H bond (in kJ mol^{-1})

$\frac{1}{2} \times (712) + \frac{3}{2} \times (436) - 3X = -46.0$

$3X = 1056$ kJ mol^{-1}

So, $X = 352$ kJ mol^{-1}.

36.(B) **I.** $C_2H_2 + \frac{5}{2}O_2 \rightarrow 2CO_2 + H_2O$; $\Delta H = -310$ kcal

II. $C + O_2 \rightarrow CO_2$; $\Delta H = -94$ kcal

III. $H_2 + \frac{1}{2}O_2 \rightarrow H_2O$; $\Delta H = -68$ kcal

Eq. (II) × 2 + Eq. (III) – Eq. (I)

$2C + H_2 \longrightarrow C_2H_2$; $\Delta H = \left[\{(-94\times2)+(-68)\}-(-310)\right] = +54$ kcal

37.(C) $2NH_3(g) \rightarrow N_2(g) + 3H_2(g)$

$\Delta H_r = -\left(2 \times \text{enthalpy of formation of } NH_3\right) = -\left(2 \times -46\right) = 92\,kJ$

38.(B) $\Delta H = \Delta U + \Delta n_g RT$

ΔH = enthalpy change (at constant pressure)

ΔU = internal energy change (at constant volume) given reaction is exothermic)

Δn_g = mole of (gaseous products – gaseous reactants)

$= -ve$ Thus, $\Delta H < \Delta U$

39.(C) The relationship between ΔH and ΔU is $\Delta H = \Delta U + \Delta nRT$

40.(C) The bond energy is the energy required to break one mole of similar bonds.

41.(C) The heat of reaction for an ideal gas, at constant pressure and constant volume are related as

$\therefore$ $\Delta H = \Delta E + \Delta nRT$ $\because$ $q_p = q_v + \Delta nRT$

42.(C) $\Delta H_{reaction} = \sum \Delta H_f$ product $- \sum \Delta H_f$ reactant

ΔH_f of a single element O_2, H_2, Cl_2 is considered equal to zero.

$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$

$\Delta H_{reaction} = \Delta H_f(CO_2) + 2\Delta H_f \cdot (H_2O) - \Delta H_f(CH_4) - 2\Delta H_f(O_2)$

$= -400 + 2(-280) - \left(-70\right) - 0 = -890$ kJ mol^{-1}

43.(D) Given, $S(rh) + \frac{3}{2}O_2(g) \longrightarrow SO_3(g)$; $\Delta H = -2x$ kJ mol^{-1} **...(i)**

$SO_2(g) + \frac{1}{2}O_2(g) \longrightarrow SO_3(g)$; $\Delta H = -y$ kJ mol^{-1} **...(ii)**

$S(rh) + O_2(g) \longrightarrow SO_2(g)$; $\Delta H = ?$

Subtract Equation (ii) from Equation (i)

$S(rh) + \frac{3}{2}O_2(g) \longrightarrow SO_3(g)$; $\Delta H = -2x$ kJ mol^{-1} **...(i)**

$SO_2(g) + \frac{1}{2}O_2(g) \longrightarrow SO_3(g)$; $\Delta H = -y$ kJ mol^{-1} **...(ii)**

$S(rh) + \left(\frac{3}{2} - \frac{1}{2}\right)O_2(g) \longrightarrow SO_2(g)$; $\Delta H = \left(-2x + y\right)$kJ mol^{-1}

$S(rh) + O_2(g) \longrightarrow SO_2(g)$; $\Delta H = \left(y - 2x\right)$kJ mol^{-1}

44.(A) Given, $H_2O(g) + C(g) \longrightarrow CO(g) + H_2(g)$; $\Delta H = +131$kJ **...(i)**

$CO(g) + \frac{1}{2}O_2(g) \longrightarrow CO_2(g)$; $\Delta H = -282$ kJ **...(ii)**

$H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(g)$; $\Delta H = -242$ kJ **...(iii)**

$C(g) + O_2(g) \longrightarrow CO_2(g)$; $\Delta H = ?$ **...(iv)**

On adding Equations (i), (ii) and (iii), we get Equation (iv)

$$H_2O(g) + C(g) \longrightarrow CO(g) + H_2(g); \quad \Delta H = +131 \text{ kJ}$$

$$CO(g) + \frac{1}{2}O_2(g) \longrightarrow CO_2(g); \quad \Delta H = -282 \text{ kJ}$$

$$\frac{H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(g); \quad \Delta H = -242 \text{ kJ}}{C(g) + O_2(g) \longrightarrow CO_2(g) \quad \Delta H = (131 - 282 - 242)\text{KJ}} = -393 \text{ kJ}$$

45.(B) $A(g) + 2B(g) \longrightarrow 2C(g) + 3D(g)$

$\Delta n = 5 - 3 = 2$

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

$19000 = \Delta E + 2 \times 2 \times 300$

$\Delta E = 17800 \text{ cal} = 17.8 \text{ kcal}$