Thermo-chemistry[chapter-7]

SHORT QUESTION AND ANSWERS

Q.1   Define the following terms and give three examples of each.

Ans.

(i)   System:

 The substance which is under experiment or under observation is called as system.

 Examples:

(i)   Pb(NO3)2 in decomposition of Pb(NO3)2.

(ii)   Zn and CuSO4 solution, the reaction mixture in the vessel.

(iii)   CaCO3 in thermal decomposition of CaCO3

(ii)   Surroundings:

 Everything around the system which is not a part of system is called surroundings.

 For example

 During the reaction between Zn and CuSO4 solution vessel and air etc are surroundings.

(iii)   State function:

 A macroscopic property of a system which has some definite value for initial and final state and independent of the path followed e.g.

 (i) Pressure (ii) Temperature (iii) Internal energy.

Note: Heat is not a state function.

Q.2   Describe the units of energy.

Ans.

 Mostly Joule and calorie are used for the measurement of energy.

Calorie:

 The amount of energy required to raise the temperature of one gram of water from 14.5oC to 15.5oC is called one calorie.

Joule:

 It is SI unit of energy and defined as energy expanded when a force of one Newton moves a body through one meter in the direction in which force is applied.
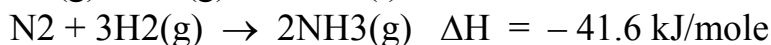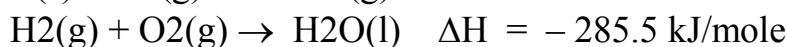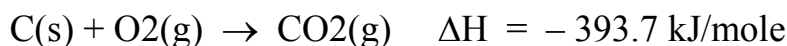
 Joule = Force x distance

        1J =   1 N x 1 m

Q.3   What are rhermochemical reactions?

Ans:

        Exothermic reaction:

 Those thermochemical reactions in which heat is evolved as a result of reaction are called as exothermic reactions.

$$C(s) + O2(g) \rightarrow CO2(g) \quad \Delta H = -393.7 \text{ kJ/mole}$$
$$H2(g) + O2(g) \rightarrow H2O(l) \quad \Delta H = -285.5 \text{ kJ/mole}$$
$$N2 + 3H2(g) \rightarrow 2NH3(g) \quad \Delta H = -41.6 \text{ kJ/mole}$$

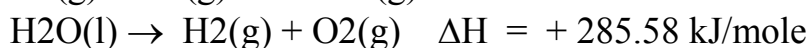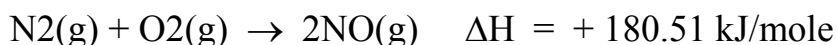Endothermic reactions:

Those thermochemical reactions in which heat is absorbed as a result of reaction are called as endothermic reaction.

$$N2(g) + O2(g) \rightarrow 2NO(g) \quad \Delta H = +180.51 \text{ kJ/mole}$$
$$H2O(l) \rightarrow H2(g) + O2(g) \quad \Delta H = +285.58 \text{ kJ/mole}$$
$$H2(g) + I2(g) \rightarrow 2H \, I \quad \Delta H = +52.96 \text{ kJ/mole}$$

Q.4 Differentiate between internal energy and enthalpy of a system?

Ans:

Internal energy:

The total of all kinds of K.E and P.E of all the particles of a system is called as internal energy. It is denoted by "E' e.g., kinetic energy may be in the form of translation, vibrational and rotational motion and potential energy is intermolecular and intramolecular forces of attraction.It is a sate function of system.

E= K.E + P.E

Enthalpy of the system:

The total heat contents of a system and denoted by H. The increase in the internal energy of a system plus work done is called as enthalpy i.e.
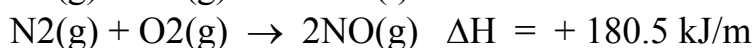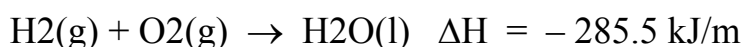
$$H = E + Pv$$

Q.5 Define the followings:
(i) Enthalpy of reaction
(ii) Enthalpy of neutralization
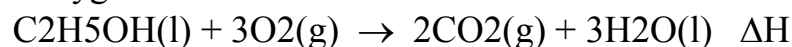**(iii) Enthalpy of combustion**

Ans:

Standard Enthalpy of reaction:

The enthalpy change when no. of moles of reactants as indicated by the balanced chemical equation react completely together to give the products under the standard conditions.
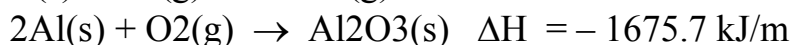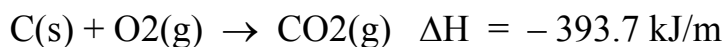
$$H2(g) + O2(g) \rightarrow H2O(l) \quad \Delta H = -285.5 \text{ kJ/m}$$
$$N2(g) + O2(g) \rightarrow 2NO(g) \quad \Delta H = +180.5 \text{ kJ/m}$$

Standard Enthalpy of Combustion:

The enthalpy change when one mole of a substance is completely burnt in excess of oxygen under standard conditions.

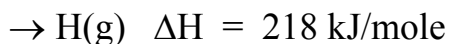`
$$C2H5OH(l) + 3O2(g) \rightarrow 2CO2(g) + 3H2O(l) \quad \Delta H$$
$$= -1368 \text{ kJ/m}$$

$$C(s) + O2(g) \rightarrow CO2(g) \quad \Delta H = -393.7 \text{ kJ/m}$$
$$2Al(s) + O2(g) \rightarrow Al2O3(s) \quad \Delta H = -1675.7 \text{ kJ/m}$$
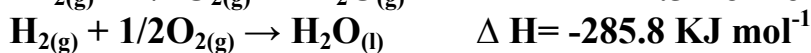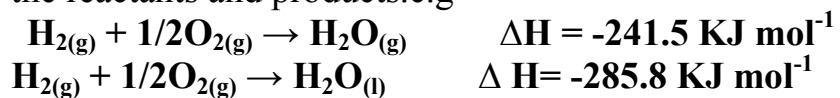
Standard enthalpy of atomization:

The enthalpy change when one mole of gaseous atoms are formed from the elements under the standard conditions is called enthalpy of atomization.

$\rightarrow$ H(g) ΔH = 218 kJ/mole

Cl2 $\rightarrow$ Cl(g) ΔH = + 121 kJ/mole

Q.6   Why it is essential to mention the physical states of reactants and products in a thermochemical equation?

Ans.

The heat of reaction depends upon the physical states of the reactants and products, heat of reaction is different in different physical states therefore, while writing a thermochemical equation it is essential to mention the physical states of the reactants and products.e.g

$H_{2(g)} + 1/2O_{2(g)} \rightarrow H_2O_{(g)}$ $\Delta H = -241.5 \text{ KJ mol}^{-1}$

$H_{2(g)} + 1/2O_{2(g)} \rightarrow H_2O_{(l)}$ $\Delta H = -285.8 \text{ KJ mol}^{-1}$

Q.7   Differentiate between spontaneous and non–spontaneous reaction?

Ans.

The process which takes place on its own without any outside help and moves from a non–equilibrium state to equilibrium state is called spontaneous or natural process. It is real, unidirectional and irreversible e.g. water flows from higher level to low leve, reaction between acid and base etc.

There are certain reactions which need energy to start but once they start they proceed their own for example burning of candle.

The process which does not take place on its own and does not occur in nature is called as non–spontaneous. It is reverse of the spontaneous process i.e. pumping of water uphill, flow of heat from colder to hotter region etc.

Q.8   **Prove that change in enthalpy is equal to heat of reaction? / prove that qp =ΔH?**

Ans.

We know that enthalpy is equal to the internal energy plus product of pressure and volume.

$H = E + Pv$

According to first law of the thermodynamics

$q = \Delta E + w$

At constant pressure $w = P\Delta v$

$q_p = \Delta E + P\Delta v$

$q_p = (E2 - E1) + (v2 - v1)p$

$q_p = E2 - E1 + Pv2 - Pv1$

$q_p = (E2 + Pv2) - (E1 + Pv1)$ $H2 = E2 + Pv2$

qp = H2 – H1   H1 = E1 + Pv1
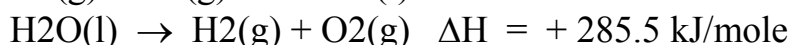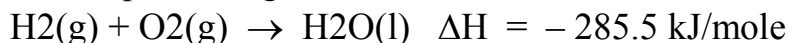
qp = ΔH

This shows that enthalpy change is equal to amount of heat absorbed at constant pressure.

Q.9   Briefly explain laws of thermochemistry.

Ans.   First law of thermochemistry:

The enthalpy of formation of a compound to the enthalpy of decomposition of that compound. e.g.

$H2(g) + O2(g) \rightarrow H2O(l)$ ΔH = – 285.5 kJ/mole

$H2O(l) \rightarrow H2(g) + O2(g)$ ΔH = + 285.5 kJ/mole

Second law of thermochemistry (Hess's law):

The amount of heat evolved or absorbed in a chemical reaction is same whether the reaction takes place in one or several steps. e.g. single step process.
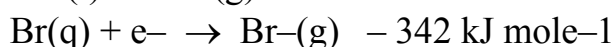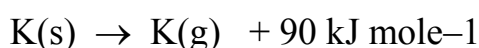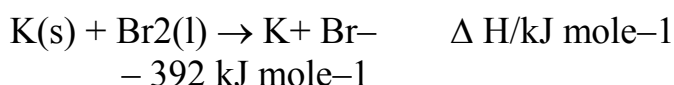
$C(s) + O2(g) \rightarrow CO2(g)$ ΔH1 = – 393.7 kJ

Two steps process:

$C(s) + O2(g) \rightarrow CO(g)$ ΔH2 = – 110.7kJ
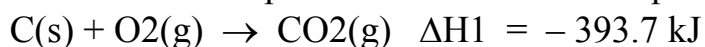
ΔH1 = ΔH2 + ΔH3

– 393.7 = – 110.7 – 283

– 393.7 = – 393.7

Q.10   Draw a complete, fully labeled Born–Haber cycle for the formation of KBr.

Reactions:

        $K(s) + Br2(l) \rightarrow K+ Br–$ Δ H/kJ mole–1

                – 392 kJ mole–1

        $K(s) \rightarrow K(g)$ + 90 kJ mole–1

        $K(g) \rightarrow K+ e–$ + 420 kJ mole–1

        $Br2(l) \rightarrow Br(g)$ +112 kJ mole–1

        $Br(q) + e– \rightarrow Br–(g)$ – 342 kJ mole–1

Solution:

The heat of formation of KBr is equal to sum of all the enthalpies.

ΔHf = – 392 kJ mole–1   (Heat of formation)

ΔHs = + 90 kJ mole–1   (Heat of sublimation)

ΔHi = + 420 kJ mole–1 (Ionization pot of k)

ΔHD/2 = + 112 kJ mole–1   (Dissociation energy of Br2)

ΔHe = – 342 kJ mole   (electron affinity)

ΔHl = ?   (Lattice energy of kBr)

ΔHf = ΔHat + ΔHi + ΔHd + ΔHe + ΔHl

– 392 = + 90 + 420 + 112 + (– 342) + ΔHl

ΔHl = – 672 kJ mole–1

Q.11   Heat is evolved in exothermic reactions and absorbed in endothermic relations.

Ans.

  When bond formation energy is greater than the bond breaking energy then the excess of energy is evolved making the reaction exothermic. When the bond breaking energy is   greater than the bond formation energy then the difference of energy is supplied from surrounding making the reaction is endothermic.

Q.12   How would you explain that change in enthalpy is a state function?

Ans.

  As $H = E + PV$

  As E, P and v are state functions as they are independent of path and depend only on the initial and final state of the system therefore enthalpy of a system is also a state function because enthalpy depends on E, P and V.

Q.13   How can you prove that $(w = -P\Delta V)$

Ans.

  We know that

  Work $=$ Force x distance

  Initial volume= V1

  Final volume = V2

  Change in volume $\Delta V = V2 - V1$

In the expansion of gases, the work is expressed in terms of pressure and change in volume. So,

          Work $=$ $P \times \Delta V$

          $=$ $w = -P\Delta V$

The work is negative because it is done by the system.

Q14.   What is difference between heat and temperature? Write a mathematical relationship between these two parameters?

Ans:

Heat: The measure of total energy of a substance is called heat. It is property of a body which flows from a body at higher temperature to a body at lower temperature. It is denoted by 'q'. It depends upon the quantity of a substance. It is measured by calorimeter. It is not a state function.

Temperature: It is measure of average K.E of the molecules in the system. It is denoted by 'T'. It is independent of the quantity of a substance. It is measured by thermometer. It is a state function.

Realtionship: $q = m \times S \times \Delta T$

Q15. **What is enthalpy of neutralization? Why enthalpy of neutralization of strong acid and base is always -57.4 KJ mol^{-1}?**

Ans:

It is the amount of heat eveolved or absorbed when one mole of H^+ ions from an acid reacts with one mole of OH^- from a base to form one mole of H_2O. Under standard conditions it is called standard enthalpy of neutralization, and it is denoted by ΔH_n.

$H^+ + OH^- \rightarrow H_2O \qquad \Delta H_n = -57.4$ KJ mol^{-1}

The heat of neutralization of strong acid or base is always -57.4 KJ mol^{-1} because strong acid or base is completely ionized and when acid and base is mixed no bond has to be broken