CBSE Class 11 Chemistry
Important Questions
Chapter 6
Thermodynamics

1 Marks Questions

1.Define a system.

Ans. A system in thermodynamics refers to that part of the universe in which observations are made.

2.Define surroundings.

Ans. The rest of the universe which might be in a position to exchange energy and matter with the system is called its surroundings.

3.State the first law of thermodynamics.

Ans. The first law of thermodynamics stales that ‘the energy of an isolated system is constant’.

4.What kind of system is the coffee held in a cup?

Ans . Coffee held in a cup is an open system because it can exchange matter (water vapors) and energy (heat) with the surroundings.

5.Give an example of an isolated system.

Ans. Coffee held in a thermos flask is an isolated system because it can neither exchange energy nor matter with the surroundings.

6.Name the different types of the system.

Ans. There are three types of system –

(i) Open system

(ii) Closed system

(iii) Isolated system.

7.What will happen to internal energy if work is done by the system?

Ans. The internal energy of the system will decrease if work is done by the system.

8.From thermodynamic point of view, to which system the animals and plants belong?

Ans. Open system.

9.How may the state of thermodynamic system be defined?

Ans. The state of thermodynamic system may be defined by specifying values of state variables like temperature, pressure, volume.

10.Define enthalpy.

Ans. It is defined as total heat content of the system.

11.Give the mathematical expression of enthalpy.

Ans. Mathematically,

H = U + pv where U is internal energy.

12.When is enthalpy change (ΔH) -

(i) positive

(ii) negative.

Ans. (i) ΔH is positive for endothermic reaction which absorbs heat from the surroundings.

(ii) ΔH is negative for exothermic reactions which evolve heat to the surroundings.

13.Give the expression for

(i)isothermal irreversible change, and

isothermal reversible change.

Ans.(i) For isothermal irreversible change Q = -w = pex (v_f - v_i)

(ii) For isothermal reversible changeq = -w = nRT ln $\frac{v_f}{v_i}$

= 2.303 nRT log $\frac{v_f}{v_i}$

14.Define Heat capacity

Ans. The heat capacity for one mole of the substance is the quantity of heat needed fo raise the temperature of one mole by one degree Celsius.

15.Define specific heat.

Ans. Specific heat /specific heat capacity is the quantity of heat required to raise the temperature of one unit mass of a substance by one degree Celsius (or one Kelvin).

16.Give the mathematical expression of heat capacity.

Ans. The mathematical expression of heat capacity

q = c x m x ΔT (c = heat capacity) when m = 1

= $C\Delta T$

where C = specific heat

m = mass

ΔT = temperature change.

17.Define reaction enthalpy.

Ans. The enthalpy change accompanying a reaction is called the reaction enthalpy $\left(\Delta_r H\right)$.

18.Define standard enthalpy.

Ans. The standard enthalpy of reaction is the enthalpy change for a reaction is the enthalpy change for a reaction when all the participating substances are in their standard states.

19.The standard heat of formation of Fe_2O_3 (s) is 824.2kJ mol^{-1} Calculate heat change for the reaction.

$4Fe(s) + 3O_2\ (g) \rightarrow 2Fe_2O_3(s)$

Ans. $\Delta H^\circ = \sum \Delta H_f^\circ(products) - \sum \Delta H_f^\circ(reactants)$

$= [2 \times \Delta H_f^0\ Fe_2O_3(s)] - [4\ \Delta H_f^0 Fe\ (s) + 3\ \Delta H_f^0 O_2(g)]$

$= 2(-824.2kJ) - [4 \times 0 + 3 \times 0]$

$= \underline{\underline{-1648.4kJ}}$

20.Define spontaneous process.

Ans. A spontaneous process is an irreversible process and may only be reversed by some external agency.

21.Define non-spontaneres process.

Ans. A process is said to be non-spontaneous if it does not occur of its own under given condition and occur only when an external force is continuously applied.

22.What is the sign of enthalpy of formation of a highly stable compound?

Ans. Negative.

23.Predict the sign of ΔS for the following reaction

$CaCo_3$ (s) $\xrightarrow{\Delta} CaO_5(s) + CO_2(g).$

Ans. ΔS is positive.

24.Two ideal gases under same pressure and temperature are allowed to mix in an isolated system – what will be sign of entropy change?

Ans. Entropy change is positive. It is because disorder or degree of freedom increase on mixing.

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2 Marks Questions

1.Change in internal energy is a state function while work is not, why?

Ans. The change in internal energy during a process depends only upon the initial and final state of the system. Therefore it is a state function. But the work is related the path followed. Therefore, it is not a state function.

2.With the help of first law of thermodynamics and H = U + pv, prove ΔH = q_p

Ans. The enthalpy is defined as

H = U + pv

For a change in the stales of system,

$\Delta H = \Delta\left(U + \text{pv}\right)$

$= \Delta U + \Delta(pv)$

$= \Delta U + p\Delta v + v\Delta p$ …………(i)

The first law of thermodynamics states that –

$\Delta U = q + w$

$= q - \Delta v$ ……………………..(ii)

From (i) and (ii),

$\Delta H = q - \cancel{\Delta} v + p\cancel{\Delta} v + v\Delta p$

= $q + V\Delta p$

When the pressure is constant,

$\Delta p = o$, then vΔp=o,

$\therefore \Delta H = q\,(at\text{ constant pressure})$

$\Delta H = qp$

3.Why is the difference between ΔH and ΔU not significant for solids or liquids?

Ans. The difference between ΔH and ΔU is not usually significant for systems consisting of only solids and / or liquids because they do not suffer any significant volume changes upon heating.

4.What is an extensive and intensive property?

Ans. Extensive property is a property whose value depends on the quantity or size of matter present in the system.

Intensive property is a property which do not depend upon the quantity or size of matter present.

5.Show that for an ideal gas Cp- Cv = R

Ans. When a gas is heated under constant pressure, the heat is required for raising the temperature of the gas and also for doing mechanical work against the external pressure during expansion.

At constant volume, the heat capacity, C is written as Cv and at constant pressure this is denoted by Cp.

we write heat q

at constant volume as qv = Cv $\Delta T = \Delta$U at constant pressure as qp = $Cp\Delta T = \Delta H$

The difference between Cp and Cv can be derived for an ideal gas as :

For a mole of an ideal gas, $\Delta H = \Delta U + \Delta(pv)$

= $\Delta U + \Delta(RT)$

= $\Delta U + R\Delta T$

$\therefore \quad \Delta H = \Delta U + R\Delta T$ _ _ _ _ _ _ _ _ (i)

On putting the values of ΔH and ΔU, we have;

$Cp\Delta T = Cv\Delta T + R\Delta T$

Cp = Cv +R

$Cp - Cv = R$

6.Show that for an ideal gas, the molar heat capacity under constant volume conditions is equal to 3/2 R.

Ans. For an ideal gas, from kinetic theory of gases, the average kinetic energy per mole (Ek) of the gas at any temperature T_k is given by $Ek = \frac{3}{2}RT$

At (T+1)k, the kinetic energy per mole (Ek^1) is Ek^1 = $\frac{3}{2}R(T+1)$

Therefore increase in the average kinetic energy of the gas for 1^0C (or 1K) rise in temperature is $\Delta\overline{E}_k = \frac{3}{2}R(T+1) - \frac{3}{2}RT = \frac{3}{2}R$

$\overline{E}_k$ by definition is to the molar heat capacity of a gas at constant volume, Cv.

$\therefore Cv = \frac{3}{2}R$

7.A 1.25g sample of octane (C_{18} H_{18}) is burnt in excess of oxygen in a bomb calorimeter. The temperature of the calorimeter rises from 294.05 to 300.78K. If heat capacity of the calorimeter is 8.93 KJ/K. find the heat transferred to calorimeter.

Ans . Mass of octane,

M = 1.250g.

= 0.00125.

Heat capacity, c = 8.93 kJ/k

Rise in temp, $\Delta T = 300.78 - 294.05$

= 6.73K

Heat transferred to calorimeter

$= m \text{ x } c \text{ x } \Delta T$

= 0.00125 x 8.93 x 6.73

= 0.075 kJ

8.Calculate the heat of combustion of ethylene (gas) to from CO_2 (gas) and H_2O (gas) at 298k and 1 atmospheric pressure. The heats of formation of CO_2, H_2O and C_2H_4 are – 393.7, - 241.8, + 52.3 kJ per mole respectively.

Ans. C_2H_4 (g) + $3O_2$(g) $\longrightarrow$ $2CO_2$(g) + $2H_2O$ (g)

$\Delta H_f(CO_2) = -393.7kJ$

$\Delta H_f(H_2O) = -241.8kJ$

$\Delta H_f(C_2H_4) = +52.3kJ$

$\Delta Hreaction = \sum \Delta H_f^o products - \sum \Delta H_f^o$ reactants

= [2 x ΔH_f^O (CO_2) + 2 x $\Delta H_f^O(H_2O)$] – $\left[\Delta H_f^O(C_2H_4) + 3 \text{ x } \Delta H_f^O(O_2)\right]$

= 2 x[(-393.7)m+2x (-241.8)] – [(523.0) + 0)]

$\left[\therefore \Delta H_f^O \text{ fot elementary substance} = 0\right]$

= [-787.4 – 483.6] -53.3

= - 1323.3 kJ.

9.Give two examples of reactions which are driven by enthalpy change.

Ans. Examples of reactions driven by enthalpy change:

The process which is highly exothermic, i.e. enthalpy change is negative and has large value but entropy change is negative is said to be driven by enthalpy change, eg.

(i) $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$

$\Delta H_f^O = -285.8 kJ\ mol^{-1}$

(ii) $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

$\Delta H^O = -92k\ J\ mol^{-1}$.

10.Will the heat released in the following two reactions be equal? Give reasons in support of your answer.

(i)H_2 **(g)** + $\frac{1}{2}O_2(g) \rightarrow H_2O(g)$

$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$

Ans. No, the heats released in the two reactions are not equal. The heat released in any reaction depends upon the reactants, products and their physical states. Here in reaction (i), the water produced is in the gaseous state whereas in reaction (ii) liquid is formed. As we know, that when water vapors condensed to from water, heat equal to the latent heat of vaporization is released. Thus, more heat is released in reaction (ii).

11.What is the relation between the enthalpy of reaction and bond enthalpy?

Ans .A chemical reaction involves the breaking of bonds in reactants and formation of new bonds in products. The heat of reaction (enthalpy change) depends on the values of the heat needed to break the bond formation .Thus

(Heat of reaction = (Heat needed to break the bonds in reactants – Heat liberated to from bonds in products).

ΔH^{O} = Bond energy in (to break the bonds) - Bond energy out (to form the bonds)

= Bond energy of reactants – Bond energy of products.

12.The reaction C (graphite) + O_2 (g) $\rightarrow$ CO_2(g) + 393.5 kJ mol^{-1} represents the formation of CO_2 and also combustion of carbon. Write the ΔH^{O} values of the two processes.

Ans.(i) The standard enthalpy of formation of CO_2 is -393.5 kJ per mole of CO_2.

That is $\Delta H_f^{O}(CO_2, g) = -393.5 kJ\ mol^{-1}$.

(ii) The standard enthalpy of combustion of carbon is - 393.5 kJ per mole of carbon i.e. $\Delta H^{O} comb(c) = -393.5\ kJ\ mol^{-1}$.

13.Explain how is enthalpy related to spontaneity of a reaction?

Ans. Majority of the exothermic reactions are spontaneous because there is decrease in energy.

Burning of a substance is a spontaneous process.

C(s) +O$_2$(g) $\rightarrow CO_2(g)$; $\Delta H = -394 \text{kJ mol}^{-1}$

Neutralisation of an acid with a base is a spontaneous reaction.

$HCl + NaOH \rightarrow NaCl + H_2O.\Delta H = -57 \text{kJ mol}^{-1}$

Many spontaneous reactions proceed with the absorption of heat. Conversion of water into water vapour is an endothermic spontaneous change. Therefore change in enthalpy is not the only criterion for deciding the spontaneity of a reaction.

14.The ΔH and ΔS for $2Ag_2O(s) \rightarrow 4Ag(s) + O_2(g)$ are given + 61.17kJ mol^{-1} and + 132 Jk^{-1}mol^{-1} respectively. Above what temperature will the reaction be spontaneous?

Ans. The reaction

2Ag$_2$O (s) $\rightarrow 4Ag(s) + O_2(g)$

Will be spontaneous when ΔG is negative.

Since ΔH is +ve and ΔS is also +ve, the relation

$$\Delta G = \Delta H - T\Delta S$$

Shows that ΔG would be –ve when,

$$\Delta H < T\Delta S$$

Or T > $\dfrac{\Delta H}{\Delta S} = \dfrac{61170 \text{ mol}^{-1}}{132 Jk^{-1} mol^{-1}} = 463.4K$

$\therefore$ The process will be spontaneous above a temperature of $\underline{\underline{463.4K}}$.

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3 Marks Questions

1.Give the relationship between ΔU and $\Delta \mathrm{H}$ for gases.

Ans. For gases the volume change is appreciable.

let V_A be the total volume of gaseous reactants, and

V_B be the total volume of gaseous product.

n_A be the number of moles of the reactant and

n_B be the number of moles of the product,

Then at constant pressure and temperature,

$p V_A = n_A RT$

$p V_B = n_B RT$

or $p V_B - pV_A = (n_B - n_A) RT$

or p $\Delta V = (\Delta n)_g \ \mathrm{RT}$

where $(\Delta n)_g = n_B n_A$ and is equal to the difference between the number of moles of gaseous products and gaseous reactants.

Substituting the value of p ΔV awe get.

$$\Delta H = \Delta U + (\Delta n)_g RT$$

$\therefore \Delta H = qp$ (heat change under constant pressure)

$\Delta U = qv$ (heat change under constant volume)

$\therefore$ for gaseous system.

$qp = qv + (\Delta n)_g RT$

2.It has been found that 221.4J is needed to heat 30g of ethanol from 15^{0}C to 18^{0}C. calculate (a) specific heat capacity, and (b) molar heat capacity of ethanol.

Ans.(a) Specific heat capacity

$$\frac{\textit{Heat}\text{ absorbed by the substance}}{\textit{Mass}\text{ of the substance x Rise in temp.}}$$

$$C = \frac{221.4j}{30g\ (18^0c - 15^0c)} = \frac{221.4}{30 \text{ x } 3}\ Jg^{-1}\ {}^0c^{-1}$$

$$= \underline{\underline{2.46J}}\ g^{-1}\ {}^0c^{-1}$$

Since 1^{0}C is equal to 1k, the specific heat capacity of ethanol = 2.46Jg^{-1} ^{0}c^{-1}.

(b) Molar heat capacity, Cm = specific heat x molar mass.

Therefore, Cm (ethanol) = 2.46 x 46

= 113.2 Jmol^{-1} ^{0}c^{-1}

The molar heat capacity of ethanol is 113.2 J mol^{-1} ^{0}c^{-1}.