DPP - Daily Practice Problems

Date : ☐ **Start Time :** ☐ **End Time :** ☐

CHEMISTRY CC06

SYLLABUS : Thermodynamics

Max. Marks : 120 **Marking Scheme :** + 4 for correct & (–1) for incorrect **Time : 60 min.**

INSTRUCTIONS : This Daily Practice Problem Sheet contains 30 MCQ's. For each question only one option is correct. Darken the correct circle/ bubble in the Response Grid provided on each page.

1. Assuming that water vapour is an ideal gas, the internal energy change (ΔU) when 1 mol of water is vapourised at 1 bar pressure and 100°C, (given : molar enthalpy of vapourisation of water at 1 bar and 373 K = 41 kJ mol^{-1} and R = 8.3 J mol^{-1} K^{-1}) will be

(a) 41.00 kJ mol^{-1} (b) 4.100 kJ mol^{-1}

(c) 3.7904 kJ mol^{-1} (d) 37.904 kJ mol^{-1}

2. In a closed insulated container, a liquid is stirred with a paddle to increase the temperature, which of the following is true?

(a) $\Delta E = W \neq 0, q = 0$ (b) $\Delta E = W = q \neq 0$

(c) $\Delta E = 0, W = q \neq 0$ (d) $W = 0, \Delta E = q \neq 0$

3. The maximum work which a system can perform at constant temperature and pressure equals

(a) $\Delta H - \Delta E$

(b) $(-\Delta G)$

(c) $q - \Delta E$

(d) $P.\Delta V - \Delta G$

4. The entropy change in the fusion of one mole of a solid melting at 27ºC (Latent heat of fusion, 2930 J mol^{-1}) is :

(a) 9.77 J K^{-1} mol^{-1}

(b) 10.73 J K^{-1} mol^{-1}

(c) 2930 J K^{-1} mol^{-1}

(d) 108.5 J K^{-1} mol^{-1}

RESPONSE GRID **1.** ⓐⓑⓒⓓ **2.** ⓐⓑⓒⓓ **3.** ⓐⓑⓒⓓ **4.** ⓐⓑⓒⓓ

5. The enthalpy of vapourisation of water from the following two equations is.

$$H_2(g)+\frac{1}{2}O_2(g) \longrightarrow H_2O\ (l),\ \Delta H=-286\,kJ$$

$$H_2\ (g)+\frac{1}{2}O_2\ (g),\rightarrow H_2O\ (g),\ \Delta H=-245.5\,kJ$$

(a) 6.02 kJ (b) 40.5 kJ
(c) 62.3 kJ (d) 1.25 kJ

6. The standard enthalpy of formation of NH_3 is -46.0 kJ mol^{-1}. If the enthalpy of formation of H_2 from its atoms is -436 kJ mol^{-1} and that of N_2 is -712 kJ mol^{-1}, the average bond enthalpy of N – H bond in NH_3 is

(a) -964 kJ mol^{-1} (b) $+352$ kJ mol^{-1}
(c) $+1056$ kJ mol^{-1} (d) -1102 kJ mol^{-1}

7. Calculate enthalpy change for the change $8S(g)$ ¾¾® $S_8(g)$, given that

$$H_2S_2(g)\longrightarrow 2H(g)+2S(g), \Delta H = 239.0\ kcal\,mol^{-1}\ ;$$

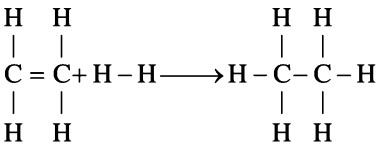
$$H_2S(g)\longrightarrow 2H(g)+S(g), \Delta H = 175.0\,kcal\,mol^{-1}$$

(a) $+512.0$ k cal (b) -512.0 k cal
(c) 508.0 k cal (d) -508.0 k cal

8. From the following bond energies:

H – H bond energy: 431.37 kJ mol^{-1}
C = C bond energy: 606.10 kJ mol^{-1}
C – C bond energy: 336.49 kJ mol^{-1}
C – H bond energy: 410.50 kJ mol^{-1}

Enthalpy for the reaction,



$$H_2C=CH_2 + H-H \longrightarrow H_3C-CH_3$$

will be:

(a) -243.6 kJ mol^{-1} (b) -120.0 kJ mol^{-1}
(c) 553.0 kJ mol^{-1} (d) 1523.6 kJ mol^{-1}

9. On the basis of the following thermochemical data : $[\Delta_f H°(H^+_{(aq)} = 0)]$

$$H_2O(l) \rightarrow H^+(aq)+OH^-(aq); \Delta H = 57.32\,kJ$$

$$H_2(g)+\frac{1}{2}O_2(g)\longrightarrow H_2O(\ell);\ \Delta H=-286.20\,kJ$$

The value of enthalpy of formation of OH^- ion at 25° C is:

(a) –228.88 kJ (b) +228.88 kJ
(c) –343.52 kJ (d) –22.88 kJ

10. For a particular reversible reaction at temperature T, ΔH and ΔS were found to be both +ve. If T_e is the temperature at equilibrium, the reaction would be spontaneous when

(a) $T_e > T$ (b) $T > T_e$
(c) T_e is 5 times T (d) $T = T_e$

11. Consider the following spontaneous reaction

$$2C_8H_{18}(g)+25\,O_2(g)\rightarrow 16CO_2(g)+18H_2O(g)$$

The sign of ΔH, ΔS, ΔG would be respectively

(a) +, –, + (b) –, +, –
(c) +, +, – (d) –, –, –

12. The favourable conditions for a spontaneous reaction are

(a) $T\Delta S > \Delta H,\ \Delta H = +ve,\ \Delta S = +ve$
(b) $T\Delta S > \Delta H,\ \Delta H = +ve,\ \Delta S = -ve$
(c) $T\Delta S = \Delta H,\ \Delta H = -ve,\ \Delta S = -ve$
(d) $T\Delta S = \Delta H,\ \Delta H = +ve,\ \Delta S = +ve$

13. The enthalpies of the following reactions are shown below.

$$\frac{1}{2}H_2(g)+\frac{1}{2}O_2(g)\rightarrow OH(g)\ ;\ \Delta H = 42.09\ kJ\ mol^{-1}$$

$$H_2(g)\rightarrow 2H(g)\ ;\ \Delta H = 435.89\ kJ\ mol^{-1}$$

$$O_2(g)\rightarrow 2O(g)\ ;\ \Delta H = 495.05\ kJ\ mol^{-1}$$

Calculate the O — H bond energy for the hydroxyl radical.

(a) 223.18 kJ mol^{-1} (b) 423.38 kJ mol^{-1}
(c) 513.28 kJ mol^{-1} (d) 113.38 kJ mol^{-1}

14. When 5 litres of a gas mixture of methane and propane is perfectly combusted at 0°C and 1 atmosphere, 16 litre of oxygen at the same temperature and pressure is consumed. The amount of heat released from this combustion in kJ (ΔH_{comb} (CH_4) = 890 kJ mol^{-1}, ΔH_{comb} (C_3H_8) = 2220 kJ mol^{-1}) is

(a) 32 (b) 38
(c) 317 (d) 477

RESPONSE GRID					
	5. ⓐⓑⓒⓓ	6. ⓐⓑⓒⓓ	7. ⓐⓑⓒⓓ	8. ⓐⓑⓒⓓ	9. ⓐⓑⓒⓓ
	10. ⓐⓑⓒⓓ	11. ⓐⓑⓒⓓ	12. ⓐⓑⓒⓓ	13. ⓐⓑⓒⓓ	14. ⓐⓑⓒⓓ

15. The heat of combustion of carbon to CO_2 is –393.5 kJ/mol. The heat released upon formation of 35.2 g of CO_2 from carbon and oxygen gas is

(a) –315 kJ (b) +315kJ

(c) –630 kJ (d) –3.15 kJ

16. Using the data provided, calculate the multiple bond energy (kJ mol^{-1}) of a $C \equiv C$ bond in C_2H_2. That energy is (take the bond energy of a C – H bond as 350 kJ mol^{-1})

$$2C(s) + H_2(g) \longrightarrow HC \equiv CH(g); \quad \Delta H = 225 \text{ kJ mol}^{-1}$$

$$2C(s) \longrightarrow 2C(g); \Delta H = 1410 \text{ kJ mol}^{-1}$$

$$H_2(g) \longrightarrow 2H(g); \Delta H = 330 \text{ kJ mol}^{-1}$$

(a) 1165 (b) 837

(c) 865 (d) 815

17. For a given reaction, $\Delta H = 35.5$ kJ mol^{-1} and $\Delta S = 83.6$ JK^{-1} mol^{-1}. The reaction is spontaneous at : (Assume that ΔH and ΔS do not vary with tempearature)

(a) T > 425 K (b) All temperatures

(c) T > 298 K (d) T < 425 K

18. If an endothermic reaction is non-spontaneous at freezing point of water and becomes feasible at its boiling point, then

(a) ΔH is –ve, ΔS is +ve

(b) ΔH and ΔS both are +ve

(c) ΔH and ΔS both are –ve

(d) ΔH is +ve, ΔS is -ve

19. The standard enthalpies of formation of CO_2(g), H_2O(l) and glucose(s) at 25°C are –400 kJ/mol, –300 kJ/mol and –1300 kJ/mol, respectively. The standard enthalpy of combustion per gram of glucose at 25°C is

(a) +2900 kJ (b) –2900 kJ

(c) –16.11 kJ (d) +16.11 kJ

20. If at 298 K the bond energies of C — H, C — C, C = C and H — H bonds are respectively 414, 347, 615 and 435 kJ mol^{-1}, the value of enthalpy change for the reaction

$$H_2C = CH_2(g) + H_2(g) \to H_3C — CH_3(g) \text{ at } 298 \text{ K}$$

will be

(a) – 250 kJ (b) + 125 kJ

(c) – 125 kJ (d) + 250 kJ

21. Standard entropy of X_2, Y_2 and X Y_3 are 60, 40 and 50 J K^{-1} mol^{-1}, respectively. For the reaction,

$\frac{1}{2}X_2 + \frac{3}{2}Y_2 \to XY_3, \Delta H = -30 \text{kJ}$, to be at equilibrium, the temperature will be

(a) 1250 K (b) 500 K

(c) 750 K (d) 1000 K

22. The enthalpy changes for the following processes are listed below :

$Cl_2(g) \to 2Cl(g)$,	242.3 kJ mol^{-1}
$I_2(g) \to 2I(g)$,	151.0 kJ mol^{-1}
$ICl(g) \to I(g) + Cl(g)$,	211.3 kJ mol^{-1}
$I_2(s) \to I_2(g)$,	62.76 kJ mol^{-1}

Given that the standard states for iodine and chlorine are $I_2(s)$ and $Cl_2(g)$, the standard enthalpy of formation for ICl(g) is :

(a) +16.8 kJ mol^{-1}

(b) +244.8 kJ mol^{-1}

(c) –14.6 kJ mol^{-1}

(d) –16.8 kJ mol^{-1}

23. Bond enthalpies of H_2, X_2 and HX are in the ratio 2 : 1 : 2. If enthalpy of formation of HX is – 50 kJ mol^{-1}, the bond enthalpy of X_2 is

(a) 100 kJ mol^{-1} (b) 300 kJ mol^{-1}

(c) 200 kJ mol^{-1} (d) 400 kJ mol^{-1}

24. Given that:

(i) $\Delta_f H°$ of N_2O is 82 kJ mol^{-1}

(ii) Bond energies of N ≡ N, N = N, O = O and N = O are 946, 418, 498 and 607 kJ mol^{-1} respectively,

The resonance energy of N_2O is :

(a) –88kJ (b) –66kJ

(c) –62kJ (d) –44kJ

RESPONSE GRID	**15.** ⓐⓑⓒⓓ	**16.** ⓐⓑⓒⓓ	**17.** ⓐⓑⓒⓓ	**18.** ⓐⓑⓒⓓ	**19.** ⓐⓑⓒⓓ
	20. ⓐⓑⓒⓓ	**21.** ⓐⓑⓒⓓ	**22.** ⓐⓑⓒⓓ	**23.** ⓐⓑⓒⓓ	**24.** ⓐⓑⓒⓓ

25. Given that bond energies of H – H and Cl – Cl are 430 kJ mol^{-1} and 240 kJ mol^{-1} respectively and ΔH_f for HCl is – 90 kJ mol^{-1}, bond enthalpy of HCl is

(a) 380 kJ mol^{-1} (b) 425 kJ mol^{-1}

(c) 245 kJ mol^{-1} (d) 290 kJ mol^{-1}

26. Under isothermal condition for one mole of ideal gas what is the ratio of work done under reversible to irreversible process, initially held at 20 atm undergoes expansion from 1L to 2L, at 298K, under external pressure of 10 atm?

(a) 1.7 (b) 2.0

(c) 1.4 (d) 1.0

27. Temperature of 5 moles of a gas is decreased by 2K at constant pressure. Indicate the correct statement

(a) Work done by gas is = 5 R

(b) Work done over the gas is = 10 R

(c) Work done by the gas = 10 R

(d) Work done = 0

28. For complete combustion of ethanol,

$C_2H_5OH(l)+3O_2(g)\longrightarrow 2CO_2(g)+3H_2O(l)$,

the amount of heat produced as measured in bomb calorimeter, is 1364.47 kJ mol^{-1} at 25°C. Assuming ideality the enthalpy of combustion, Δ_cH, for the reaction will be: ($R = 8.314$ kJ mol^{-1})

(a) –1366.95 kJ mol^{-1} (b) –1361.95 kJ mol^{-1}

(c) –1460.95 kJ mol^{-1} (d) –1350.50 kJ mol^{-1}

29. Read the following statements carefully and choose the correct answer

(i) Expansion of a gas in vacuum ($p_{ex} = 0$) is called free expansion.

(ii) Work is done during free expansion of an ideal gas whether the process is reversible or irreversible

(iii) No work is done during free expansion of an ideal gas whether the process is reversible or irreversible

(iv) No work is done during free expansion of an ideal gas when the process is reversible

(a) Only statement (iii) is correct

(b) Statements (i) and (iii) are correct

(c) Statements (ii) and (iv) are correct

(d) Statements (i) and (iv) are correct

30. The (S°) of the following substances are:

CH_4 (g) 186.2 JK^{-1} mol^{-1}

O_2 (g) 205.2 JK^{-1} mol^{-1}

CO_2 (g) 213.6 JK^{-1} mol^{-1}

H_2O (g) 69.9.JK^{-1} mol^{-1}

The entropy change ($\Delta S°$) for the reaction

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$ is:

(a) – 312.5 J K^{-1} mol^{-1}

(b) – 242.8 J K^{-1} mol^{-1}

(c) – 108.1 J K^{-1} mol^{-1}

(d) – 37.6 J K^{-1} mol^{-1}

RESPONSE GRID	**25.** ⓐⓑⓒⓓ **26.** ⓐⓑⓒⓓ **27.** ⓐⓑⓒⓓ **28.** ⓐⓑⓒⓓ **29.** ⓐⓑⓒⓓ **30.** ⓐⓑⓒⓓ

DAILY PRACTICE PROBLEM DPP CHAPTERWISE 6 - CHEMISTRY			
Total Questions	30	Total Marks	120
Attempted		Correct	
Incorrect		Net Score	
Cut-off Score	36	Qualifying Score	52
Success Gap = Net Score – Qualifying Score			
Net Score = (Correct × 4) – (Incorrect × 1)			

1. **(d)** Given $\Delta H = 41 \text{ kJ mol}^{-1} = 41000 \text{ J mol}^{-1}$
$T = 100°C = 273 + 100 = 373 \text{ K}$
$n = 1$
$\Delta U = \Delta H - \Delta nRT = 41000 - (1 \times 8.314 \times 373)$
$= 37898.88 \text{ J mol}^{-1} \simeq 37.9 \text{ kJmol}^{-1}$

2. **(a)** Internal energy is dependent upon temperature and according to first law of thermodynamics total energy of an isolated system remains same, i.e., in a system of constant mass, energy can neither be created nor destroyed by any physical or chemical change but can be transformed from one form to another

$\Delta E = q + W$

For closed insulated container, $q = 0$, so, $\Delta E = +W$, as work is done by the system

3. **(b)** Maximum work done by a system at constant pressure and constant temperature is $-\Delta G$. It can be derived as follows :
From first law of thermodynamics,
$q = \Delta E + w_{exp.} + w_{non\ exp.}$
At constant pressure, $w_{exp} = P\Delta V$
$q = \Delta E + P\Delta V + w_{non\text{-}exp.}$
$q = \Delta H + w_{non\ exp.}$ .... (1)
For the reversible change at const temp.,

$\Delta S = \frac{q_{rev}}{T}$ or $q_{rev} = T\Delta S$

so from equ. (1), $T\Delta S = \Delta H + w_{non\ exp.}$
$\Rightarrow \Delta H - T\Delta S = -w_{non\ exp.}$ ...... (2)
For a change taking place under constant pressure and constant temperature,
$\Delta H - T\Delta S = \Delta G$
so from equ (2)
$\Delta G = -w_{non\ exp}.$
(it means $w_{non\ exp.}$ or w_{useful})
or $w_{max} = -\Delta G$

4. **(a)** $\Delta S = \frac{\text{Latent heat of fusion}}{\text{Melting point}} = \frac{\Delta H}{T}$

$= \frac{2930}{300} \text{ J K}^{-1} \text{ mol}^{-1} = 9.77 \text{ J K}^{-1} \text{ mol}^{-1}$

5. **(b)** Desired equation is $H_2O(l) \rightarrow H_2O(g)$, $\Delta H = ?$
(Equation II - Equation I)
$\Delta H = -245.5 \text{ kJ} - (-286 \text{ kJ}) = 40.5 \text{ kJ}$

6. **(b)** $N_2 + 3H_2 \longrightarrow 2NH_3 \quad \Delta H = 2 \times -46.0 \text{ kJ mol}^{-1}$
Let x be the bond enthalpy of N – H bond then
[Note : Enthalpy of formation or bond formation enthalpy is given which is negative but the given reaction involves bond breaking hence values should be taken as positive.]
$\Delta H = \Sigma$ Bond energies of reactants – Σ Bond energies of products
$2 \times -46 = 712 + 3 \times (436) - 6x$
$-92 = 2020 - 6x$
$6x = 2020 + 92$
$6x = 2112$
$x = +352 \text{ kJ/mol}$

7. **(b)** $\Delta H_{S-S} + 2\Delta H_{H-S} = 239 \quad 2\Delta H_{H-S} = 175$
Hence, $\Delta H_{S-S} = 239 - 175 = 64 \text{ kcal mol}^{-1}$
Then, ΔH for $8S_{(g)} \rightarrow S_{8(g)}$ is $8 \times (-64) = -512 \text{k cal}$

8. **(b)** Enthalpy of reaction
$= B.E_{(Reactant)} - B.E_{(Product)}$
$= \left[B.E_{(C=C)} + 4B.E._{(C-H)} + B.E._{(H-H)}\right]$
$-\left[B.E._{(C-C)} + 6B.E._{(C-H)}\right]$
$= [606.1 + (4 \times 410.5) + 431.37)] - [336.49 + (6 \times 410.5)]$
$= -120.0 \text{ kJ mol}^{-1}$

9. **(a)** Given, for reaction
(i) $H_2O(\ell) \longrightarrow H^+(aq.) + OH^-(aq.)$;
$\Delta H_r = 57.32 \text{ kJ}$
(ii) $H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(\ell)$;
$\Delta H_r = -286.20 \text{ kJ}$
For reaction (i)
$\Delta H_r = \Delta H°_f(H^+.aq) + \Delta H°_f(OH^-.aq) -$
$\Delta H°_f(H_2O, \ell)$
$57.32 = 0 + \Delta H°_f(OH^-, aq) - \Delta H°_f(H_2O, \ell)$ ...(iii)
For reaction (ii)
$\Delta H_r = \Delta H°_f(H_2O, \ell) -$
$\Delta H°_f(H_2, g) - \frac{1}{2}\Delta H°_f(O_2, g)$
$-286.20 = \Delta H°_f(H_2O, \ell)$
On replacing this value in eq. (iii) we have
$57.32 = \Delta H°_f(OH^-, aq) - (-286.20)$
$\Delta H°_f(OH^-, aq) = -286.20 + 57.32 = -228.88 \text{ kJ}$

10. **(b)** At equilibrium $\Delta G = 0$
Hence, $\Delta G = \Delta H - T_e\Delta S = 0$
$\therefore \Delta H = T_e\Delta S$ or $T_e = \frac{DH}{DS}$
For a spontaneous reaction
ΔG must be negative which is possible only if $\Delta H - T\Delta S < 0$
$\therefore \Delta H < T\Delta S$
or $T > \frac{DH}{DS}$; $T_e < T$

11. **(b)** $2C_8H_{18(g)} + 25O_{2(g)} \longrightarrow 16CO_{2(g)} + 18H_2O_{(g)}$

The given reaction is a combustion reaction. Hence this will be an exothermic reaction.

i.e. $\Delta H = -ve$

Further $\Delta n = +\ ve$ i.e. $\Delta S = +ve$

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

12. **(a)** $\Delta G = \Delta H - T\Delta S$, $\Delta H + ve$, ΔS is $+$ ve; $T\Delta S > \Delta H$ for spontaneous process. It will make ΔG, $-ve$

13. **(b)** We have to calculate the enthalpy of the reaction

$OH(g) \to O(g) + H(g)$

From the given reactions, this can be obtained as follows.

$$-\left[\frac{1}{2}H_2(g)+\frac{1}{2}O_2(g)\to OH(g)\right];\ \Delta H = -42.09\ kJ\ mol^{-1}$$

$$+\frac{1}{2}\ [H_2(g)\to 2H(g)];\ \Delta H = \frac{1}{2}\times 435.89\ kJ\ mol^{-1}$$

$$+\frac{1}{2}\ [O_2(g)\to 2O(g)];\ \Delta H = \frac{1}{2}\times 495.05\ kJ\ mol^{-1}$$

Add

$$\overline{OH(g)\to H(g)+O(g)}\ ;\ \Delta H = \overline{423.38\ kJ\ mol^{-1}}$$

14. **(c)** $\underset{x}{CH_4} + \underset{2x}{2O_2} \longrightarrow CO_2 + 2H_2O$

$\underset{(5-x)}{C_3H_8} + \underset{5(5-x)}{5O_2} \longrightarrow 3CO_2 + 4H_2O$

$2x + 5(5-x) = 16$

$\Rightarrow x = 3L$

$\therefore$ Heat released

$$= \frac{3}{22.4}\times 890 + \frac{2}{22.4}\times 2220 = 317$$

15. **(a)** $\underset{12g}{C} + O_2 \to \underset{44g}{CO_2} + 393.5\ kJ/mol$

44g is formed from 12g of carbon

35.2g is formed from $\frac{12\times 35.2}{44}$ g of C

$= 9.6\ g$ of $C = 9.6/12 = 0.8$ mole

1 mole release heat 393.5 kJ

0.8 mole release heat $= 393.5\times 0.8$

$= 314.8\ kJ \approx 315\ kJ$

16. **(d)** (i) $2C(s) + H_2(g) \longrightarrow H-C\equiv C-H(g)$

$\Delta H = 225\ kJ\ mol^{-1}$

(ii) $2C(s) \longrightarrow 2C(g)$ $\Delta H = 1410\ kJ\ mol^{-1}$

$$C(s) \longrightarrow C(g)\ \Delta H = \frac{1410}{2} = 705\ kJmol^{-1}$$

(iii) $H_2(g) \longrightarrow 2H(g)$ $\Delta H = 330\ kJ\ mol^{-1}$

From equation (i) :

$$225 = \left[2\times \Delta H_{C(s)\longrightarrow C(g)} + 1\times BE_{H-H}\right]$$

$$-[2\times BE_{C-H} + 1\times BE_{C\equiv C}]$$

$225 = [1410 + 1\times 330] - [2\times 350 + 1\times BE_{C\equiv C}]$

$225 = [1410 + 330] - [700 + BE_{C\equiv C}]$

$225 = 1740 - 700 - BE_{C\equiv C}$

$BE_{C\equiv C} = 815\ kJ\ mol^{-1}$

17. **(a)** Given ΔH 35.5 kJ mol^{-1}

$\Delta S = 83.6\ JK^{-1}\ mol^{-1}$

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

For a reaction to be spontaneous, $\Delta G = -ve$

i.e., $\Delta H < T\Delta S$

$$\therefore\ T > \frac{\Delta H}{\Delta S} = \frac{35.5\times 10^3\ Jmol^{-1}}{83.6\ JK^{-1}}$$

So, the given reaction will be spontaneous at $T > 425\ K$

18. **(b)** $\Delta G = \Delta H - T\Delta S$

Since $\Delta G = \Delta H - T\Delta S$ for an endothermic reaction, $\Delta H = +ve$ and at low temperature $\Delta S = +$ ve

Hence $\Delta G = (+)\ \Delta H - T\ (+)\Delta S$

and if $T\Delta S < \Delta H$ (at low temp)

$\Delta G = +ve$ (non spontaneous)

But at high temperature, reaction becomes spontaneous i.e. $\Delta G = -ve$.

because at higher temperature $T\Delta S > \Delta H$.

19. **(c)** The standard enthalpy of the combustion of glucose can be calculated by the eqn.

$C_6H_{12}O_6(s) + 6O_2(g) \to 6CO_2(g) + 6H_2O(l)$

$\Delta H_C = 6\times \Delta H_f(CO_2) + 6\times \Delta H_f(H_2O) - \Delta H_f[C_6H_{12}O_6]$

$\Delta H^\circ = 6(-400) + 6(-300) - (-1300)$

$\Delta H^\circ = -2900\ kJ/mol$

For one gram of glucose, enthalpy of combustion

$$\Delta H^\circ = -\frac{2900}{180} = -16.11\ kJ/gm$$

20. **(c)** $CH_2 = CH_2\ (g) + H_2\ (g) \to CH_3 - CH_3$

Enthalpy change = Bond energy of reactants – Bond energy of products.

$\Delta H = 1(C=C) + 4(C-H) + 1(H-H) - 1(C-C) - 6(C-H)$

$= 1\ (C=C) + 1\ (H-H) - 1\ (C-C) - 2\ (C-H)$

$= 615 + 435 - 347 - 2\times 414 = 1050 - 1175 = -125\ kJ.$

21. **(c)** For a reaction to be at equilibrium $\Delta G = 0$. Since $\Delta G = \Delta H - T\Delta S$ so at equilibrium $\Delta H - T\Delta S = 0$ or $\Delta H = T\Delta S$

For the reaction

$$\frac{1}{2}X_2 + \frac{3}{2}Y_2 \longrightarrow XY_3;\ \ \Delta H = -30kJ\ \text{(given)}$$

Calculating ΔS for the above reaction, we get

$$\Delta S = 50 - \left[\frac{1}{2}\times 60 + \frac{3}{2}\times 40\right]JK^{-1}$$

$= 50 - (30 + 60)\ JK^{-1} = -40\ JK^{-1}$

At equilibrium, $T\Delta S = \Delta H$ $[\because \Delta G = 0]$

$\therefore\ T\times(-40) = -30\times 1000\ [\because 1kJ = 1000J]$

$$\text{or}\quad T = \frac{-30\times 1000}{-40}\ \text{or}\ 750\ K$$

22. **(a)** $I_2(s) + Cl_2(g) \longrightarrow 2ICl(g)$

$$\Delta_r H = [\Delta H(I_2(s) \rightarrow I_2(g)) + \Delta H_{I-I} + \Delta H_{Cl-Cl}] - [\Delta H_{I-Cl}]$$
$$= 151.0 + 242.3 + 62.76 - 2 \times 211.3 = 33.46$$

$$D_f H°(ICl) = \frac{33.46}{2} = 16.73 \text{ kJ/mol}$$

23. **(a)** $\frac{1}{2}H_2 + \frac{1}{2}X_2 \longrightarrow HX$

Let the bond enthalpy of $X - X$ bond be x.

$$\Delta H_f \text{ (HX)} = -50 = \frac{1}{2}\Delta H_{H-H} + \frac{1}{2}\Delta H_{X-X} - \Delta H_{H-X}$$

$$= \frac{1}{2}2x + \frac{1}{2}x - 2x = \frac{-x}{2}$$

$$\therefore \ x = 50 \times 2 = 100 \text{ kJ mol}^{-1}$$

24. **(a)** $N_2(g) + \frac{1}{2}O_2 \rightarrow N_2O(g)$

$$N \equiv N(g) + \frac{1}{2}(O = O) \rightarrow \underset{..}{\overset{\bar{..}}{N}} = \overset{+}{N} = \underset{..}{\overset{..}{O}}(g)$$

$\Delta H_f{}^o$ = [Energy required for breaking of bonds]
−[Energy released for forming of bonds]

$$= (\Delta H_{N \equiv N} + \frac{1}{2}\Delta H_{O=O} - (\Delta H_{N=N} + \Delta H_{N=O})$$
$$= (946 + \frac{1}{2} \times 498) - (418 + 607) = 170 \text{ kJ mol}^{-1}$$

Resonance energy $= 170 - 82 = 88$ kJ mol^{-1}

25. **(b)** $\frac{1}{2}H_2 + \frac{1}{2}Cl_2 \longrightarrow HCl$

$$\Delta H_{HCl} = \sum \text{B.E. of reactant} - \sum \text{B.E. of products}$$

$$-90 = \frac{1}{2} \times 430 + \frac{1}{2} + \frac{1}{2} \times 240 - \text{B.E. of HCl}$$
$$\therefore \text{ B.E. of HCl} = 215 + 120 + 90$$
$$= 425 \text{ kJ mol}^{-1}$$

26. **(a)** $-W_{irreversible} = P_{ext}(V_2 - V_1)$
$= 10$ atm (2L – 1L)
$= 10$ atm – L

$$-W_{reversible} = \int_{V_1}^{V_2} P_{ex} dv$$

$$= 2.303 \text{ nRT} \log \frac{V_2}{V_1}$$

$$= 1 \times 2.303 \times 0.0821 \text{ atm–L/K/mol} \times \log \frac{2}{1}$$

$= 16.96$ atm–L

$$\frac{W_{reversible}}{W_{irreversible}} = \frac{16.96}{10.00} = 1.69 \approx 1.7$$

27. **(b)** For 5 moles of gas at temperature T,

$PV_1 = 5RT$

For 5 moles of gas at temperature T – 2,

$PV_2 = 5R(T - 2)$

$\therefore \ P(V_2 - V_1) = 5R(T - 2 - T);$

$P\Delta V = -10R,$

$-P\Delta V = 10R$

When ΔV is negative, W is + ve.

28. **(a)** $C_2H_5OH(\ell) + 3O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(\ell)$

Bomb calorimeter gives ΔU of the reaction

Given, $\Delta U = -1364.47$ kJ mol^{-1}

$\Delta n_g = -1$

$$\Delta H = \Delta U + \Delta n_g RT = -1364.47 - \frac{1 \times 8.314 \times 298}{1000}$$

$= -1366.93$ kJ mol^{-1}

29. **(b)**

30. **(b)** $\Delta S° = S°_{CO_2} + 2 \times S°_{H_2O} - \left(S°_{CH_4} + 2 \times S°_{O_2}\right)$
$= (213.6 + 2 \times 69.9) - (186.2 + 2 \times 205.2)$
$= -242.8$ J K^{-1} mol^{-1}.