1. The process with negative entropy change is
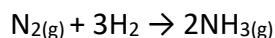(1) dissolution of iodine in water
(2) sublimation of dry ice
(3) synthesis of ammonia from N_2 and H_2
(4) dissociation of $CaSO_{4(s)}$ to $CaO_{(s)}$ and $SO_{3(g)}$.

Solution:
$N_{2(g)} + 3H_2 \rightarrow 2NH_{3(g)}$
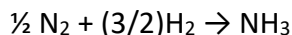$\Delta s = 2 - 4 = -2 < 0$
So entropy is negative.
Hence option (4) is the answer.

2. The standard enthalpy of formation of NH_3 is -46 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
(1) -964 kJ mol⁻¹
(2) +352 kJ mol⁻¹
(3) +1056 kJ mol⁻¹
(4) -1102 kJ mol⁻¹

Solution:
$½ N_2 + (3/2)H_2 \rightarrow NH_3$
$(\Delta H_f) NH_3 = [½ B.E N_2 + (3/2) B.E H_2 - 3 B.E N-H]$
$-46 = [½ 712 + (3/2) 436 - 3 B.E N-H]$
$-46 = 356 + 654 - 3 B.E N-H$
$3 B.E N-H = 1056$
$B.E N-H = 1056/3 = 352$ kJ mol⁻¹
Hence option (2) is the answer.

3. The enthalpy change for a reaction does not depend upon
(1) use of different reactants for the same product
(2) the nature of intermediate reaction steps
(3) the differences in initial or final temperature of involved substances
(4) the physical states of reactants and products

Solution:
The enthalpy change for a reaction does not depend upon the nature of intermediate reaction steps.
Hence option (2) is the answer.

4. Identify the correct statement regarding a spontaneous process
(1) For a spontaneous process in an isolated system, the change in entropy is positive
(2) Endothermic processes are never spontaneous

(3) Exothermic processes are always spontaneous
(4) Lowering of energy in the reaction process is the only criterion for spontaneity

Solution:
For a spontaneous process in an isolated sys-tem, the change in entropy is positive.
Hence option (1) is the answer.

5. 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   (1) $T_e > T$

(2) $T > T_e$
(3) T_e is 5 times T
(4) $T_e = T$

Solution:
At equilibrium, ΔG = 0.
ΔG = ΔH - TΔS
For a reaction to be spontaneous ΔG should be negative.
So $T > T_e$
Hence option (2) is the answer.

6. The entropy change involved in the isothermal reversible expansion of 2 moles of an ideal gas from a volume of 10 dm^3 to a volume of 100 dm^3 at 27°C is
(1) 32.3 J $mol^{-1}K^{-1}$
(2) 42.3 J $mol^{-1}K^{-1}$
(3) 38.3 J $mol^{-1}K^{-1}$
(4) 35.8 J $mol^{-1}K^{-1}$

Solution:
Given V_2 = 100
V_1 = 10
 n = 2
ΔS = 2.303nR log (V_1/V_2)
= 2.303 × 2 × 8.314 × log (100/10)
= 2.303 × 2 × 8.314 × log 10
= 38.29 J mol^{-1} K^{-1}
Hence option (3) is the answer.

7. A piston filled with 0.04 mol of an ideal gas expands reversibly from 50.0 mL to 375 mL at a constant temperature of 37^0C. As it does so, it absorbs 208J of heat. The values of q and w for the process will be   (R = 8.314J/molK and In 7.5 = 2.01)
(1) q = + 208J, w = − 208 J
(2) q = − 208J, w = − 208 J

(3) q = − 208J, w = + 208 J
(4) q = + 208J, w = + 208 J

Solution:
ΔE = q+w
The process is isothermal reversible expansion. So ΔE = 0
Hence q = -w
Since heat is absorbed, q = +208 J
So w = -208 J
Hence option (1) is the answer.

8. During compression of a spring, the work done is 10 kJ and 2 kJ escaped to the surrounding as heat. The change in internal energy ΔU (in kJ) is
(1) -8
(2) 12
(3) 8
(4) -12

Solution:
Given w = 10 kJ
q = –2 kJ
According to first law of thermodynamics, **ΔU = q + w**
= –2 + 10 = 8 kJ
Hence option (3) is the answer.

9. An ideal gas undergoes isothermal expansion at constant pressure. During the process
(1) enthalpy increases but entropy decreases
(2) enthalpy remains constant but entropy increases
(3) enthalpy decreases but entropy increases
(4) both enthalpy and entropy remain constant.

Solution:
During isothermal expansion at constant pressure, $\Delta H = n C_p \Delta T = 0$
$\Delta S = nRln(V_f/V_i) > 0$
Hence option (2) is the answer.

10. Assuming that water vapour is an ideal gas, the internal energy change (U) when 1 mol of water is vaporized 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
(1) 41.00 kJ mol⁻¹
(2) 4.100 kJ mol⁻¹
(3) 3.7904 kJ mol⁻¹
(4) 37.904 kJ mol⁻¹

Solution:
Given ΔH = 41000
T = 373 K
ΔU = ΔH-ΔnRT
= 41000 - 1× 8.314× 373
= 41000 - 3101.122
= 37898.878 J mol^{-1}
= 37.9 kJ mol^{-1}
Hence option (4) is the answer.

11. A gas undergoes change from state A to state B. In this process, the heat absorbed and work done by the gas is 5 J and 8 J, respectively. Now gas is brought back to A by another process during which 3 J of heat is evolved. In this reverse process of B to A
(1) 10 J of the work will be done by the surrounding on gas
(2) 10 J of the work will be done by the gas
(3) 6 J of the work will be done by the surrounding on gas
(4) 6 J of the work will be done by the gas.

Solution:
From A to B:
Given q = + 5 J
w = - 8 J (work done by the system)
According to first law of thermodynamics,
U = q + w
= 5 -8
= -3 J
From B to A:
U = 3 J
(As internal energy is state function and does not depend on path)
q = -3 J (heat evolved),
w = U-q
= + 3-(-3)
= + 6 J
Therefore 6 J work will be done by the surrounding on gas.
Hence option (3) is the answer.

12. An ideal gas is allowed to expand from 1 L to 10 L against a constant external pressure of 1 bar. The work done in kJ is
(1) +10.0
(2) -0.9
(3) -2.0
(4) -9.0

Solution:
Given P_{ext} = 1 bar
V_1 = 1 L
V_2 = 10 L
w = $-P_{ext}(V_2-V_1)$
= -1(10-1)
= -9 bar L
= -900 J
= -0.9 kJ
Hence option (2) is the answer.

13. Among the following, the set of parameters that represents path function, is
(A) q + w
(B) q
(C) w
(D) H – TS
(1) (B) and (C)
(2) (A) and (D)
(3) (B), (C) and (D)
(4) (A), (B) and (C)

Solution:
ΔU = q+w
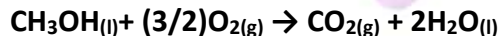q $\Rightarrow$ path function
w $\Rightarrow$ path function
ΔU $\Rightarrow$ state function
H -TS = G $\Rightarrow$ state function
Hence option (1) is the answer.

14. In a fuel cell methanol is used as fuel and oxygen gas is used as an oxidizer. The reaction is
$CH_3OH_{(l)}+ (3/2)O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(l)}$
At 298 K standard Gibb's energies of formation for CH_3OH(*l*), H_2O (*l*) and CO_2 (g) are
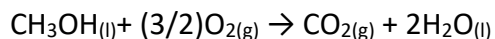**-166.2, -237.2 and -394.4 kJ mol^{-1} respectively. If the standard enthalpy of combustion of methanol is**
–726 kJmol^{-1}, the efficiency of the fuel cell will be
(1) 90%
(2) 97%
(3) 80%
(4) 87%

Solution:
$CH_3OH_{(l)}+ (3/2)O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(l)}$
Given ΔG_f^0 CH_3OH (*l*) = -166.2 kJ mol^{-1}

ΔG_f^0 H_2O (*l*) = -394.4 kJ mol^{-1}

$\Delta G = \Sigma \Delta G_f^0$ products - $\Sigma \Delta G_f^0$ reactants

= -394.4 -2(237.2) + 166.2

= -702.6 kJ mol^{-1}

The efficiency of fuel cell = ($\Delta G/ \Delta H$)× 100

= (702.6/726)× 100

= 97%

Hence option (2) is the answer.

15. Consider the reaction :

$4NO_2$ (g) + O_2 (g) → $2N_2O_5$ (g). Δ_rH = -111kJ

If N_2O_5 (s) is formed instead of N_2O_5 (g) in the above reaction, the Δ_rH value will be:-

(given, ΔH of sublimation for N_2O_5 is 54 kJ mol^{-1})

(1) –165 kJ

(2) +54 kJ

(3) +219 kJ

(4) –219 kJ

Solution:

Given $\Delta H_{reaction}$ = -111 kJ

ΔH_{sub} = 54 kJ

$\Delta H_f + \Delta H_{sub} = \Delta H_{reaction}$ [Hess's law]

$\Delta H_f = \Delta H_{reaction} - \Delta H_{sub}$

= -111 -54

= -165 kJ

Hence option (1) is the answer.