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Q-1 - 11035616

Adiabatic expansion of an ideal gas is accompanied by

- (A) Increase in temperature
- (B) Decrease in  $\Delta S$
- (C) Decrease in  $\Delta U$
- (D) No change in any one of the above properties

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Q-2 - 11035617

For a cyclic process, which of the following is true?

(A)  $\Delta S = 0$

(B)  $\Delta U = 0$

(C)  $\Delta H = 0$

(D)  $\Delta G = 0$

---

### SOLUTION:

For cyclic process,  $\Delta H = \Delta U = 0$

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Q-3 - 11035618

One mole of a gas is heated at constant pressure to raise its temperature by  $1\text{C}$ . The work done in joules is

(A)  $-4.3$

(B)  $-8.314$

(C) – 16.62

(D) Unpredictable

SOLUTION:

$$\begin{aligned}w &= - nRD\eta T \text{ or } w \\&= - p\Delta V = - 1 \\&\times 8.314 \times 1 = \\&- 8.314 J\end{aligned}$$

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Q-4 - 11035619

What work is to be done on  $2\text{mol}$  of a perfect gas at  $27^\circ\text{C}$  it is

compressed reversibly and isothermally from a pressure of

$$1.01 \times 10^5 \text{ Nm}^{-2} \rightarrow 5.05$$

$$\times 10^6 \text{ Nm}^{-2}$$

?

SOLUTION:

For reversible process:

$$w_{rev} = - 2.303nRT \log 10$$

$$\cdot \frac{P_1}{P_2}$$

$$= - 2.303 \times 2$$

$$\times 8.314 \times 300 \log_{10}$$

$$\cdot \frac{1.01 \times 10^5}{5.05 \times 10^6}$$

$$= + 1.9518 \times 10^4 J$$

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Q-5 - 11035648

5mol of an ideal gas at 27C expands isothermally and reversibly

from a volume of 6L to 60L. The work done in kJ is

(A) - 14.7

(B) - 28.72

(C) +28.72

(D) -56.72

**SOLUTION:**

$$w =$$

$$- 2.303nRT \log \frac{V_2}{V_1}$$

$$= - 2.303 \times 5$$

$$\times 8.314 \times 300 \log \frac{60}{6}$$

$$= - 28.72 \text{ kJ}$$

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Q-6 - 11035649

10mol of an ideal gas confined to a volume of 10L is released into atmosphere at 300K where the pressure is 1bar. The work done by the gas is

$$(R = 0.083 \text{ L bar K}^{-1} \text{ mol}^{-1})$$

(A)  $249 \text{ L bar}$

(B)  $259 \text{ L bar}$

(C)  $239 \text{ L bar}$

(D)  $220 \text{ L bar}$

---

SOLUTION:

Initial volume,  $V_1 = 10 \text{ L}$

$$\begin{aligned} V_2(\text{final}) &= \frac{nRT}{P} \\ &= \frac{10 \times 0.083 \times 300}{1} \\ &= 249 \text{ L} \end{aligned}$$

$$\begin{aligned} W &= P\Delta V = 1 \\ &\times (249 - 10) = 239 \text{ L} \end{aligned}$$

bar

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Q-7 - 11035651

A system absorbs  $20\text{kJ}$  heat and also does  $10\text{kJ}$  of work. The net internal energy of the system

- (A) Increases by  $10\text{kJ}$
- (B) Decreases by  $10\text{kJ}$
- (C) Increases by  $30\text{kJ}$
- (D) Decreases by  $30\text{kJ}$

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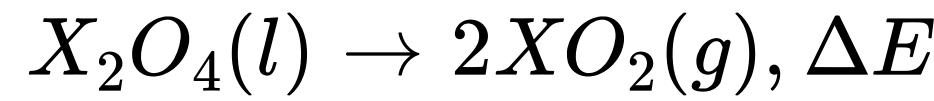
SOLUTION:

$$\Delta U = q + w = 20 - 10 = 10\text{kJ}$$

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Q-8 - 11035965

For the reaction,



$$= 2.1Kcal$$

,

$\Delta S = 20\text{cal}/K$  at  $300K$ . Hence  $\Delta G$  is

(A)  $2.7\text{kcal}$

(B)  $-2.7\text{kcal}$

(C)  $9.3\text{kcal}$

(D)  $-9.3\text{kcal}$

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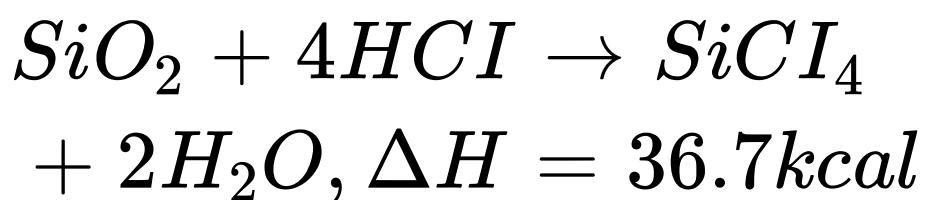
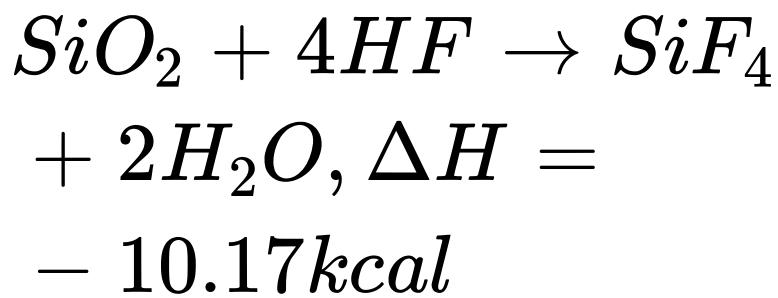
SOLUTION:

$$\begin{aligned}\Delta H &= \Delta U + \Delta n_g RT \\ &= 2.1 \times 2 \times 0.02 \\ &\times 300 = 3.3\text{kcal}\end{aligned}$$

$$\begin{aligned}\Delta G &= \Delta H - T\Delta S \\ &= 3.3 \times -300 \\ &\times (0.02) = -2.7\text{kcal}\end{aligned}$$

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For the given reactions



It may be concluded that

- (A)  $HF$  will attack  $SiO_2$  and  $HCl$  will not
- (B)  $HCl$  will attack  $SiO_2$  and  $HF$  will not
- (C)  $HF$  and  $HCl$  both attack  $SiO_2$
- (D) None attacks  $SiO_2$

---

SOLUTION:

$$\begin{aligned} \Delta G &= \Delta H - T\Delta S = \\ &- ve \end{aligned}$$

All exothermic reactions are spontaneous, hence  $HF$  will attack  $SiO_2$ .

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Q-10 - 11035968

$$\Delta G = \Delta H - T\Delta S \text{ and } \Delta G = \Delta H + T \left[ \frac{d(\Delta G)}{dT} \right]_p ,$$

then  $\left( \frac{dE_{cell}}{dT} \right)$  is

(A)  $\frac{\Delta H}{nF}$

(B)  $\frac{\Delta G}{nF}$

(C)  $\frac{\Delta S}{nF}$

(D)  $-\frac{\Delta S}{nF}$

SOLUTION:

On comparison:

$$\Delta S = - \left[ \frac{d(\Delta G)}{dT} \right]$$
$$\Delta S = - \frac{d(-nFE)}{dT}$$
$$= nF \left( \frac{dF}{dT} \right)$$

[Here E is the  
 $E_{mf}$  of cell]

$$\therefore \left( \frac{dE}{dT} \right) = \frac{\Delta S}{nF}$$

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Q-11 - 11035970

Which of the following is the intensive property?

- (A) Boiling point
- (B) Refractive index

(C) Molarity

(D) Volume

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Q-12 - 11035971

$1gH_2$  gas  $STP$  is expanded so that the volume is doubled. Hence,

work done is

(A)  $22.4L - atm$

(B)  $5.6L - atm$

(C)  $11.2L - atm$

(D)  $44.8L - atm$

**SOLUTION:**

$$V_1 (\text{volume of } 1gH_2)$$

$$= 11.2L \text{ at } STP$$

$V_2$ (volume of  $1gH_2$ )

$$= 22.4L$$

Itbr.

$$\therefore W = P\Delta V$$

$$= 11.2Latm$$

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Q-13 - 11035974

The standard heat of formation of  $AI$  is  $-837.8 kJmol^{-1}$  at  $25C$ .

If  $AI$  reacts with  $O_2$  at  $25C$ , which of the following release  $250kJ$  of heat?

(A) The reaction of  $0.624mol$  of  $AI$

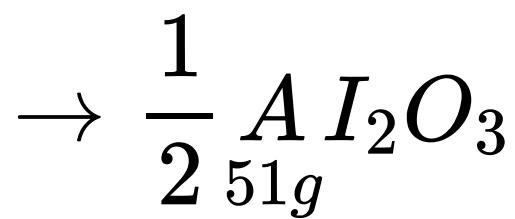
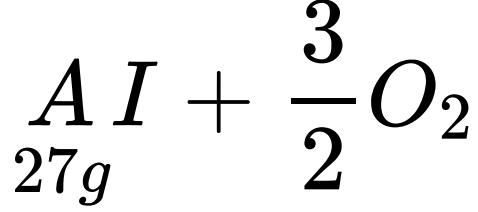
(B) The formation of  $0.624mol$  of  $AI_2O_3$

(C) The reaction of  $0.312mol$  of  $AI$

(D) The formation of  $0.150mol$  of  $AI_2O_3$

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SOLUTION:



$$\Delta H =$$

$$- 837.8 kJ mol^{-1}$$

a.

$$0.624 \text{ mol of } AI$$

$$= 837.8 \times 0.624$$

$$\text{on combustion gave } = 523 \text{ kJ.}$$

Hence false.

b. Formation of  $0.624 \text{ mol of } AI_2O_3$  gave

$$= 837.9 \times 2 \times 0.624$$

$$= 1045 \text{ kJ}$$

Hence, false.

c.  $0.312 \text{ mol of } AI$  on combustion gave  $= 261 \text{ kJ.}$

Hence, false.

d. Formation of

$0.150\text{mol of } Al_2O_3 \text{ gave}$

$$= 251.3\text{kJ}$$

. Hence, ture.

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Q-14 - 11035983

If a gas absorbs  $200\text{J}$  of heat and expands by  $500\text{cm}^3$  against a constant pressure of  $2 \times 10^5 \text{Nm}^{-2}$ , then the change in internal energy is

(A)  $-300\text{J}$

(B)  $-100\text{J}$

(C)  $+100\text{J}$

(D)  $+300\text{J}$

**SOLUTION:**

$$\Delta U = q - w = q - P$$

$$\Delta V$$

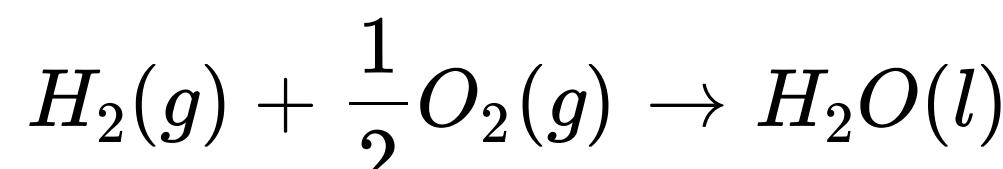
$$= 200 - 2 \times 10^5$$

$$\times 500 \times 10^{-6}$$

$$= 100J$$

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Q-15 - 11035989



$$BE(H - H) = x_1,$$

$$BE(O = O) = x_2$$

$$BE(O = H) = x_3$$

Latent heat of vaporisation of water liquid into water vapour =  $x_4$ ,

then  $\Delta_f H$  (heat of formation of liquid water) is

(A)  $x_1 + \frac{x_2}{2} - x_3 + x_4$

$$(B) 2x_3 - x_1 - \frac{x_2}{2} - x_4$$

$$(C) x_1 + \frac{x_2}{2} - 2x_3 - x_4$$

$$(D) x_1 + \frac{x_2}{2} - 2x_3 + x_4$$

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SOLUTION:

$$\Delta H = (BE)_{\text{reactant}}$$

$$- (BE)_{\text{products}}$$

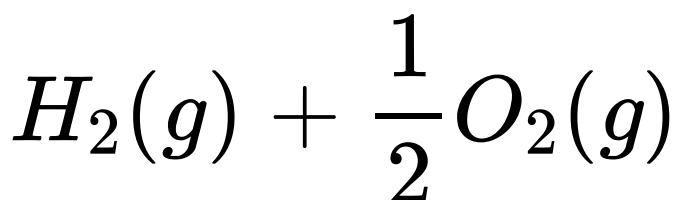
[But all the species must be in gaseous state. In product,



$$\Delta H$$

must be added.

Hence,



$$\Delta H = \left[ (BE)_{H-H} \right.$$

$$\left. + \frac{1}{2}(BE)_{O=O} \right]$$

$$= \left[ (\Delta H)_{vap} \right.$$

$$\left. + 2(BE)_{O-H} \right]$$

$$= x_1 + \frac{x_2}{2}$$

$$- [x_4 + 2x_3]$$

$$= x_1 + \frac{x_2}{2} - x_4$$

$$- 2x_3$$

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Q-16 - 11035990

If a certain mass of gas is made to undergo separately adiabatic and isothermal expansions to the same pressure, starting from the same

initial conditions of temperature and pressure, then, as compared to that of isothermal expansion, in the case of adiabatic expansion, the final

(A) Volume and temperature will be higher.

(B) Volume and temperature will be lower.

(C) Temperature will be lower but the final volume will be higher.

(D) Volume will be lower but the final temperature will be higher.

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Q-17 - 11035991

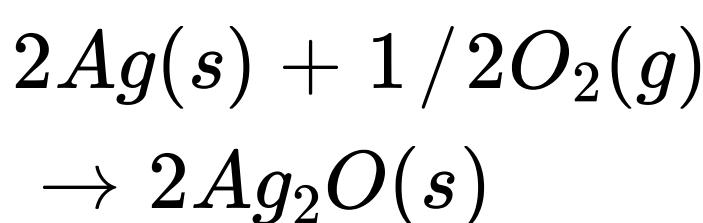
The dissolution of  $NH_4Cl$  in water is endothermic even though  $NH_4Cl$  dissolves in water spontaneously. Which one of the following best explains this behaviour?

- (A) The bonds in solid  $NH_4Cl$  are weak
- (B) The entropy-driving force causes dissolution.
- (C) Endothermic processes are energetically favourable
- (D) The dissolving process is unrelated to energy.

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Q-18 - 11035997

For the combustion reaction at 298K



Which of the following alternatives is correct?

- (A)  $\Delta H = \Delta U$
- (B)  $\Delta H > \Delta U$
- (C)  $\Delta H < \Delta U$

(D)  $\Delta H$  and  $\Delta U$  bear no relation with each other

---

SOLUTION:

$$\Delta H = \Delta U + \Delta nRT$$

$$\Delta n = n_P - n_R$$

$$\therefore \Delta n = 2 - \frac{5}{2} =$$

$$-\frac{1}{2}$$

$$\therefore \Delta H = \Delta U - \frac{1}{2}RT$$

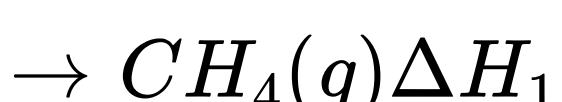
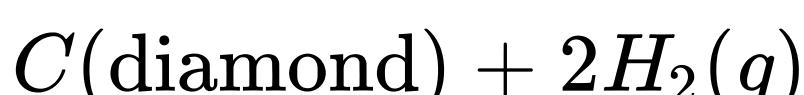
$$\therefore \Delta U = \Delta H + \frac{1}{2}RT$$

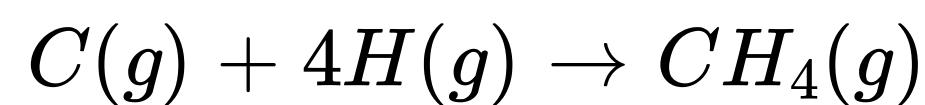
or  $\therefore \Delta U > \Delta H$

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Q-19 - 11036000

For the equations





$$\Delta H_2$$

Predict whether

(A)  $\Delta H_1 = \Delta H_2$

(B)  $\Delta H_1 > \Delta H_2$

(C)  $\Delta H_1 < \Delta H_2$

(D)

$$\Delta H_1 = \Delta H_2 +$$

$$\Delta_{vap}H(C) +$$

$$\Delta_{diss}H(H_2)$$

---

SOLUTION:

$$\Delta H_1 > \Delta H_2$$

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The enthalpy change for chemical reaction is denoted as  $\Delta H^\Theta$

and  $\Delta H^\Theta = H_P^\Theta - H_R^\Theta$ . The relation between enthalpy and

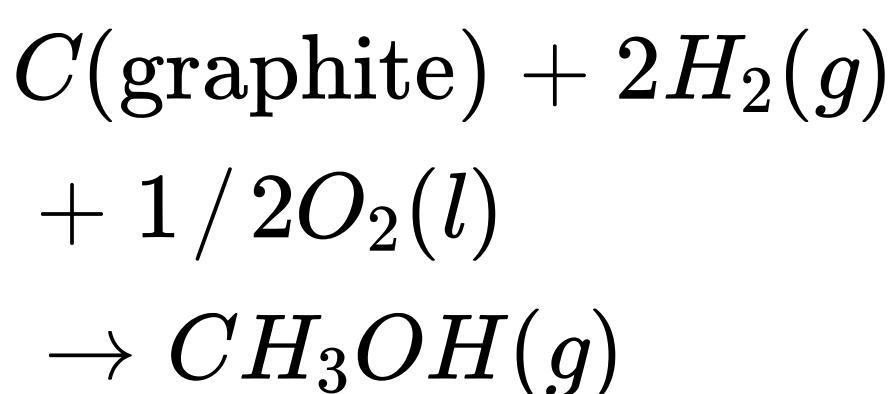
internal energy is expressed by equation:

$$\Delta H = \Delta U + \Delta nRT$$

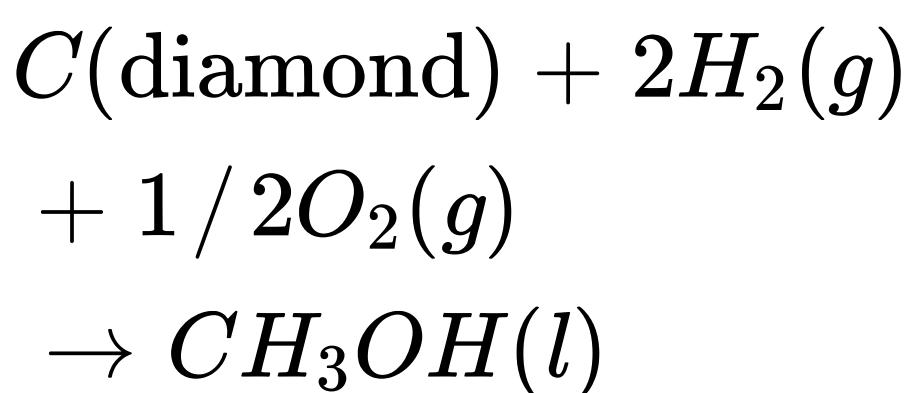
where  $\Delta U$  = change in internal energy  $\Delta n$  = change in number of moles,  $R$  = gas constant.

Which of the following equations corresponds to the definition of enthalpy of formation at 298K?

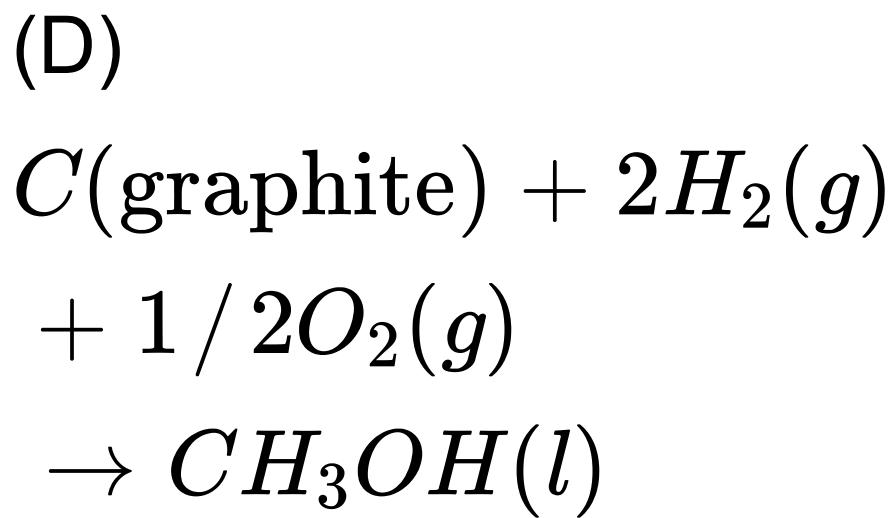
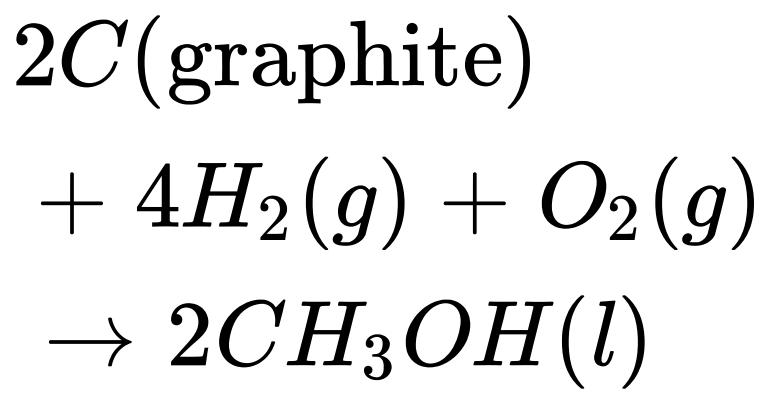
(A)



(B)



(C)



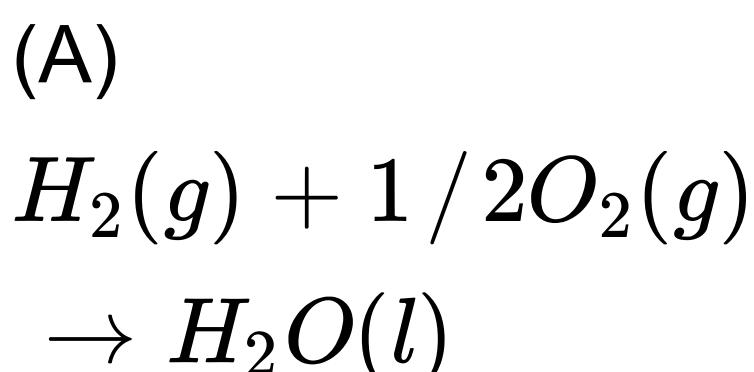
SOLUTION:

In enthalpy of formation, reactants and products must be in most stable standard state.

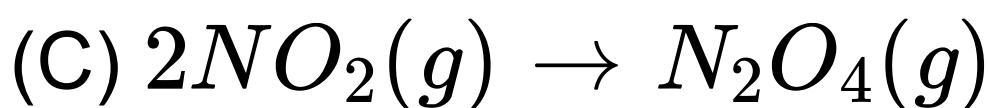
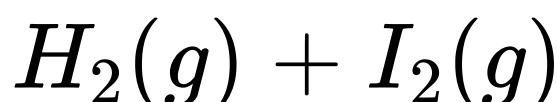
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Q-21 - 11036004

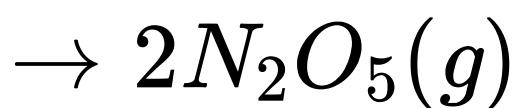
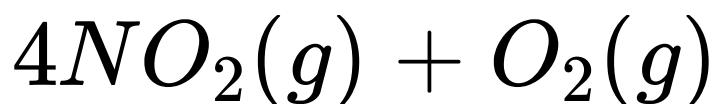
For which of the following equations, will  $\Delta H$  be equal to  $\Delta U$ ?



(B)



(D)



SOLUTION:

$$\Delta H = \Delta U + \Delta n RT$$

$$In(b), \Delta n = 0$$

$$\therefore \Delta H = \Delta U$$

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Q-22 - 11036007

The molar enthalpies of combustion of  $C_2H_2(g)$ , C(graphite) and  $H_2(g)$  are  $-1300$ ,  $-394$ , and  $-286\text{ kJ mol}^{-1}$ , respectively. The standard enthalpy of formation of  $C_2H_2(g)$  is

(A)  $-226 \text{ kJ mol}^{-1}$

(B)  $-626 \text{ kJ mol}^{-1}$

(C)  $226 \text{ kJ mol}^{-1}$

(D)  $626 \text{ kJ mol}^{-1}$

SOLUTION:

$$\Delta_f H^\Theta = (\Delta_f H_P^\Theta - \Delta_f H_R^\Theta)$$

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Q-23 - 11036008

The relationship between enthalpy and internal energy change is

(A)  $\Delta U = \Delta H + P\Delta V$

(B)  $\Delta H = \Delta U + P\Delta V$

(C)  $\Delta H = \Delta U - P\Delta V$

(D)  $P\Delta V = \Delta U + \Delta H$

---

SOLUTION:

$$\Delta H = \Delta U + P\Delta V$$

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Q-24 - 11036011

For the gaseous reaction involving the complete combustion of isobutane

(A)  $\Delta H = \Delta U$

(B)  $\Delta H > \Delta U$

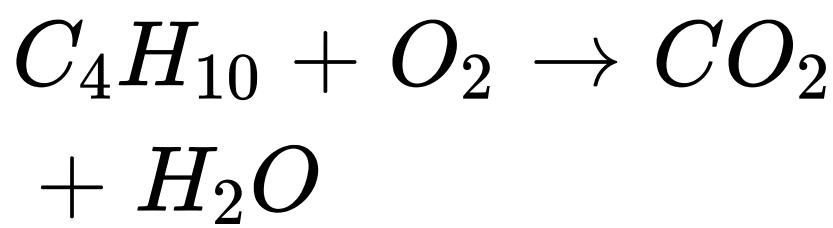
(C)  $\Delta H = \Delta U = 0$

(D)  $\Delta H < \Delta U$

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SOLUTION:

$$\Delta H > \Delta E$$



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Q-25 - 11036020

The state of equilibrium refers to

- (A)  $\Delta_{\text{total}}S > 0$
- (B)  $\Delta_{\text{total}}S < 0$
- (C)  $\Delta_{\text{total}}S = 0$
- (D) Unpredictable

**SOLUTION:**

$$\Delta S_{\text{Total}} = 0$$

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Q-26 - 11036015

For a process  $H_2O(s) \rightarrow H_2O(l)$  at 273K

(A)

$$G(\text{ice}) = G(\text{water})$$

$$= 0$$

(B)

$$G(\text{ice}) = G(\text{water}) 1$$

$$= 0$$

(C)  $G(\text{ice}) > G(\text{water})$

(D)  $G(\text{ice}) < G(\text{water})$

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SOLUTION:

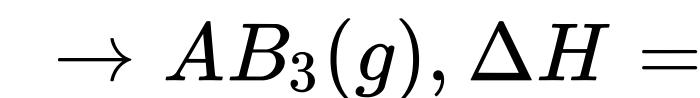
$$G(\text{ice}) = G(\text{water}) = 0$$

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Q-27 - 11036021

For hypothetical reversible reaction

$$1/2A_2(g) + 3/2B_2(g)$$



$$- 210\text{kJ}$$

if standard entropies of  $A_2$ ,  $B_2$ , and  $AB_3$  are 60, 40 and

$50\text{JK}^{-1}\text{mol}^{-1}$  respectively. The above reaction will be

equilibrium at

(A)  $400K$

(B)  $500K$

(C)  $250K$

(D)  $200K$

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SOLUTION:

At equilibrium,  $\Delta S_{\text{total}} = 0$

$$\therefore S_P^\Theta - S_R^\Theta = 0$$

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If values of  $\Delta_f H^\Theta$  of  $ICI(g)$ ,  $CI(g)$ , and  $I(g)$  are, respectively, 17.57, 121.34, and  $106.96 J mol^{-1}$ . The value of  $I - CI$  (bond energy) in  $J mol^{-1}$  is

(A)  $35.15 J mol^{-1}$

(B)  $106.69 mol^{-1}$

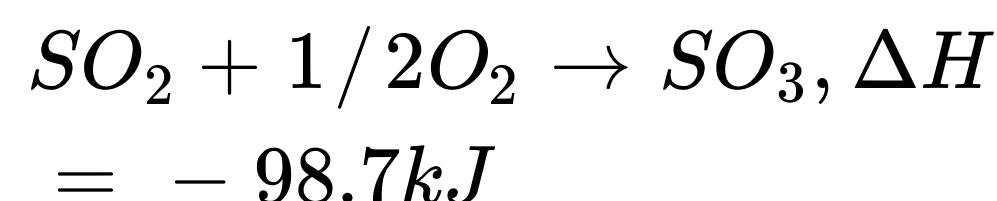
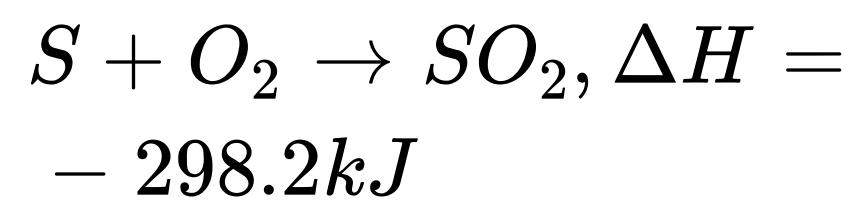
(C)  $210.73 J mol^{-1}$

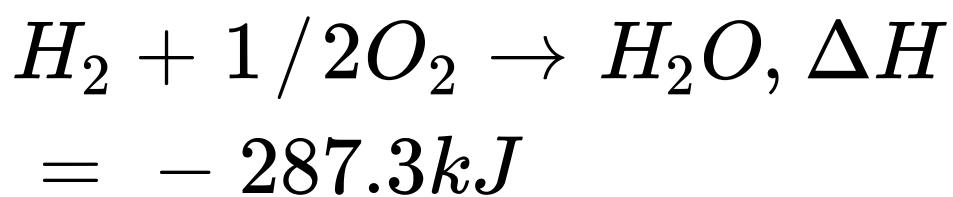
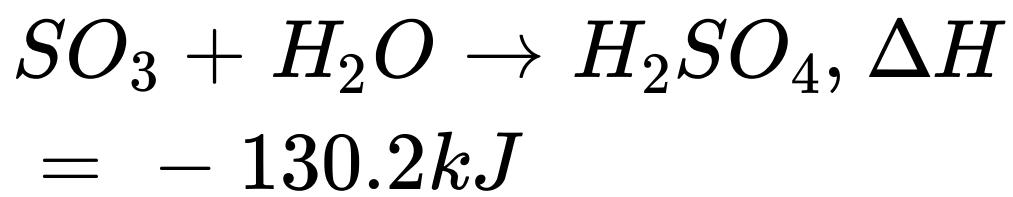
(D)  $420.0 J mol^{-1}$

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Q-29 - 11036023

If





Then the enthalpy of formation of  $H_2SO_4$  at  $298K$  is

(A)  $-814.4 \text{ kJ}$

(B)  $-650.3 \text{ kJ}$

(C)  $-320.5 \text{ kJ}$

(D)  $-233.5 \text{ kJ}$

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Q-30 - 11036030

Since the enthalpy of elements in their natural state is taken to be zero, the heat of formation ( $\Delta_f H$ ) of compounds

(A) Is always negative

(B) Is always positive

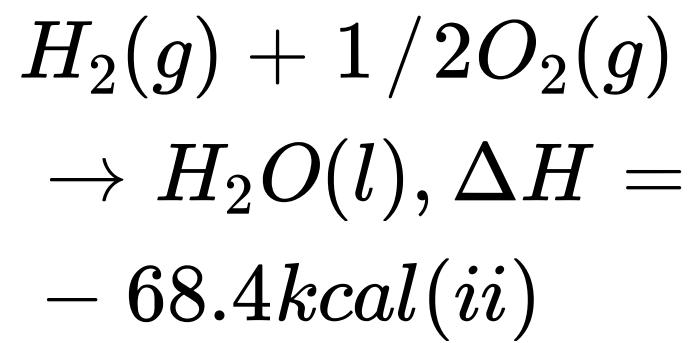
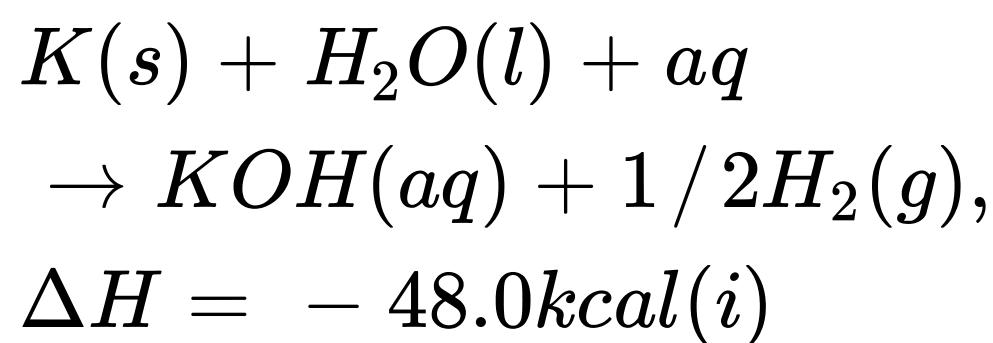
(C) May be negative or positive

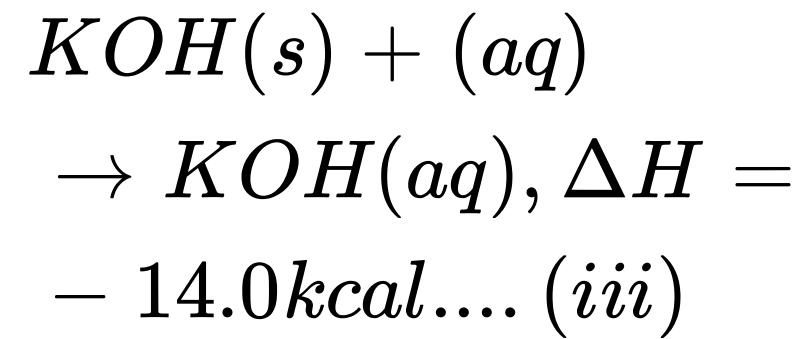
(D) Is zero

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Q-31 - 11036029

Calculate heat of formation of  $KOH(s)$  using the following equations





(A)  $-68 + 48 - 14$

(B)  $-68 - 48 + 14$

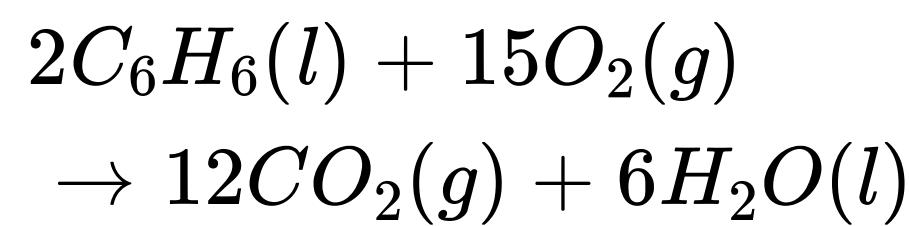
(C)  $68 - 48 + 14$

(D)  $68 + 48 + 14$

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Q-32 - 11036183

The difference between the heats of reaction at constant pressure  
and a constant volume for the reaction



at  $25^\circ C$  in  $\text{kJ}$  is

(A)  $-7.43$

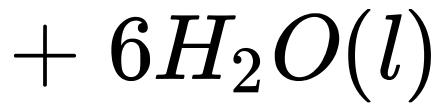
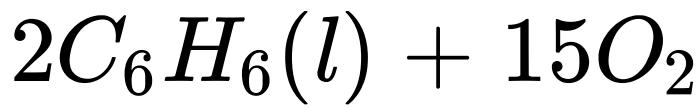
(B) + 3.72

(C) - 3.72

(D) + 7.43

---

SOLUTION:



$$\Delta H = \Delta E + \Delta n_g RT$$

$$\Delta H - \Delta E = \Delta n_g RT$$

$$\Delta n_g = 12 - 15 = -3$$

$$\Delta H - \Delta E = -3$$

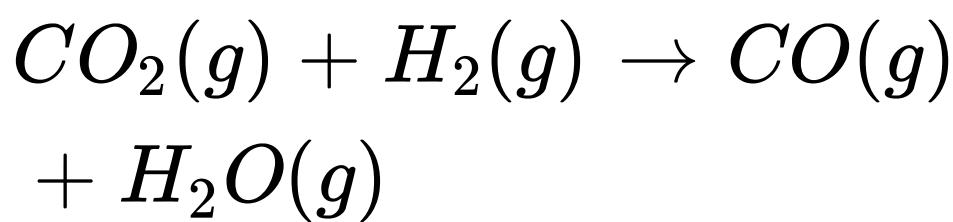
$$\times 8.314 \times 298 =$$

$$- 7432.716 J$$

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$\Delta H_1$  for  $CO_2(g)$ ,  $CO(g)$  and  $H_2O(g)$  are  $-393.5$ ,  $-110.5$  and  $-241.8 \text{ kJ mol}^{-1}$  respectively. Standard enthalpy change for the

reaction



is

(A) 524.1

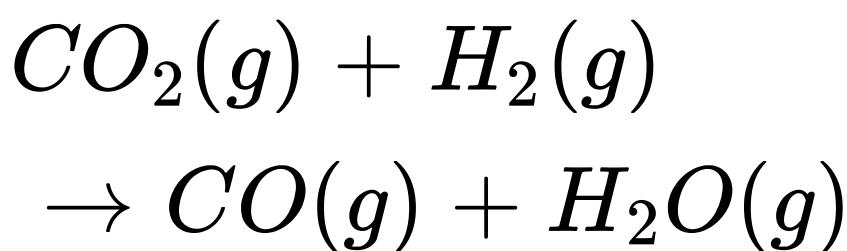
(B) +41.2

(C) -262.5

(D) -41.2

---

SOLUTION:



$$\Delta H^\Theta = \left\{ \left[ \Delta_f H^\Theta CO, (g) + \Delta_f H^\Theta H_2O, (g) \right] - \Delta_f H^\Theta CO_2, (g) \right\}$$

$$= [( - 110.5) + ( - 241.8) - ( - 393.5 + 0))]$$

$$= + 41.2 kJ$$

$$\Delta_f H^\Theta H_2, (g) = 0$$

$\Delta_f H^\Theta = 0$  in elementary state

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Q-34 - 11036188

One mole of a non-ideal gas undergoes a change of state  
 $(2.0\text{atm}, 3.0\text{L}, 95\text{K})$   
 $\rightarrow (4.0\text{atm}, 5.0\text{L}, 245\text{K})$

With a change in internal energy  $\Delta E = 30L$  atm. The change in enthalpy ( $\Delta H$ ) in the process in  $L\text{-atm}$  is

(A) 40.0

(B) 42.3

(C) 44.0

(D) Not defined because the pressure is not constant.

---

SOLUTION:

$$H = U + PV$$

$$\begin{aligned}H_1 &= U_1 + (2 \times 3) \\&= U_1 + 6\end{aligned}$$

$$\begin{aligned}H_2 &= U_2 + (4 \times 5) \\&= U_2 + 20\end{aligned}$$

$$\begin{aligned}(H_2 - H_1) \\= (U_2 - U_1) \\+ (20 - 6)\end{aligned}$$

$$\begin{aligned}\Delta H = \Delta U + 14 = 30 \\+ 14 = 44 \text{ Latm}\end{aligned}$$

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Q-35 - 52405992

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

(A)  $-1102 \text{ kJ mol}^{-1}$

(B)  $-964 \text{ kJ mol}^{-1}$

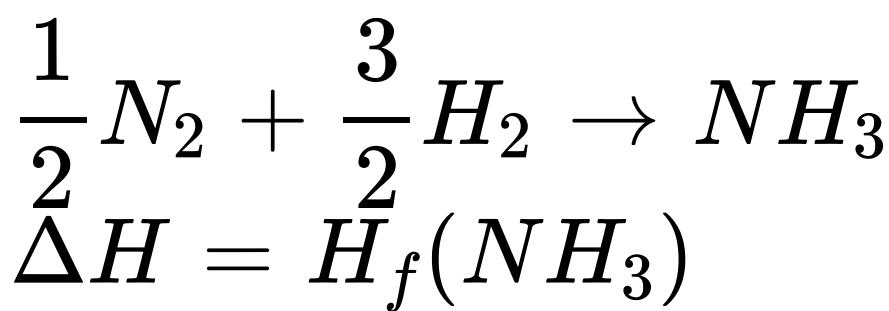
(C)  $+352 \text{ kJ mol}^{-1}$

$$(D) + 1056 kJ mol^{-1}$$

---

CORRECT ANSWER: C

SOLUTION:



$$- \frac{1}{2}H_f(N_2)$$

$$- \frac{3}{2}H_f(H_2)$$

$$-46 = H_f(NH_3)$$

$$- \frac{1}{2}(-712) - \frac{3}{2}$$

$$\times (-436)$$

$$H_f(NH_3) = -46$$

$$-356 - 654 =$$

$$-1056 KJ$$

enthalpy of formation of  $NH_3 = -1056$  KJ

Average bond enthalpy of N - H bond

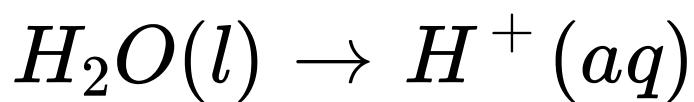
$$= \frac{1056}{2} = 352$$

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Q-36 - 52405904

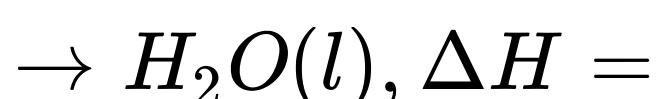
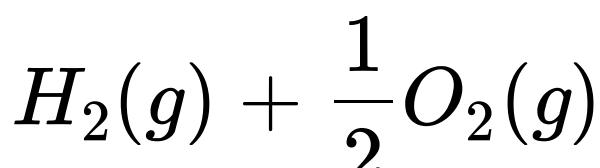
On the basis of the following thermochemical data :

$$\left( \Delta fG H_{(aq)}^+ = 0 \right)$$



$$+ OH^-(aq), \Delta H$$

$$= 57.32 kJ$$



$$- 286.20 kJ$$

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

(A)  $- 22.88 \text{ kJ}$

(B)  $- 228.88 \text{ kJ}$

(C) +228.88 kJ

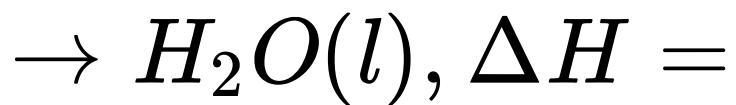
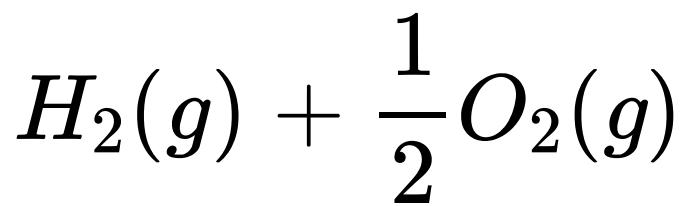
(D) -343.52 kJ

---

CORRECT ANSWER: B

---

SOLUTION:



$$- 286.20 \text{ kJ}$$

$$\Delta H_r = \Delta H_f(H_2O, l)$$

$$- \Delta H_f(H_2, g) - \frac{1}{2}$$

$$\Delta H_f(O_2, g)$$

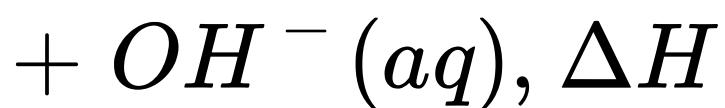
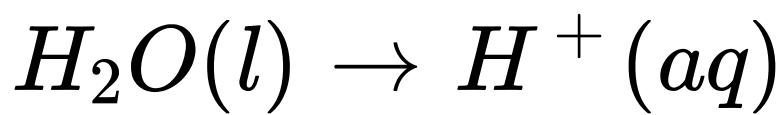
$$- 286.20 =$$

$$\Delta H_f(H_2O(l))$$

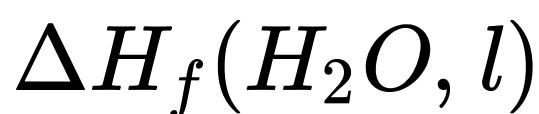
so,

$$\Delta H_f(H_2O, l) =$$

$$- 286.20$$



$$= 57.32 \text{ kJ}$$



$$57.32 = 0 +$$



$$- (- 286.20)$$



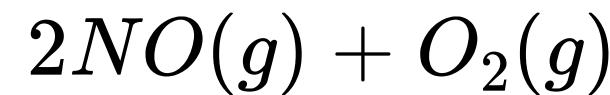
$$= 57.32 - 286.20 =$$

$$- 228.88 \text{ kJ}$$

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Q-37 - 34966062

The following reaction is performed at 298K



The standard free energy of formation of  $NO(g)$  is 86.6 KJ/mol at 298 K. What is the standard free energy formation of  $NO_2(g)$  at 298 ? ( $K_p = 1.6 \times 10^{12}$ )

(A)  $R(298) \ln (1.6 \times 10^{12}) - 86600$

(B)  $86600 + R(298) \ln (1.6 \times 10^{12})$

(C)

86600

$$-\frac{\ln(1.6 \times 10^{12})}{R(298)}$$

(D)

$$0.5 [2 \times 86,600$$

$$- R(298) \ln(1.6 \times 10^{12})]$$

---

CORRECT ANSWER: D

---

## SOLUTION:

$$2\Delta G_f(NO_2) - [2\Delta G_f(NO) + \Delta G_f(O_2)] = \Delta G_r = -RT\ln K_p$$

$$2\Delta G_f(NO_2) - [2 \times 86,600 + 0] = -RT\ln K_p$$

$$\Delta G_f(NO_2) = 0.5 [2 \times 86,600 - R(298) \ln(1.6 \times 10^{12})]$$

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Q-38 - 52406166

The heats of combustion of carbon and carbon monoxide are - 393.5 and -283.5 KJ mol<sup>-1</sup>, respectively. The heat of formation (in kJ) of carbon monoxide per mole is

(A) 676.5

(B) –676.5

(C) –110.5

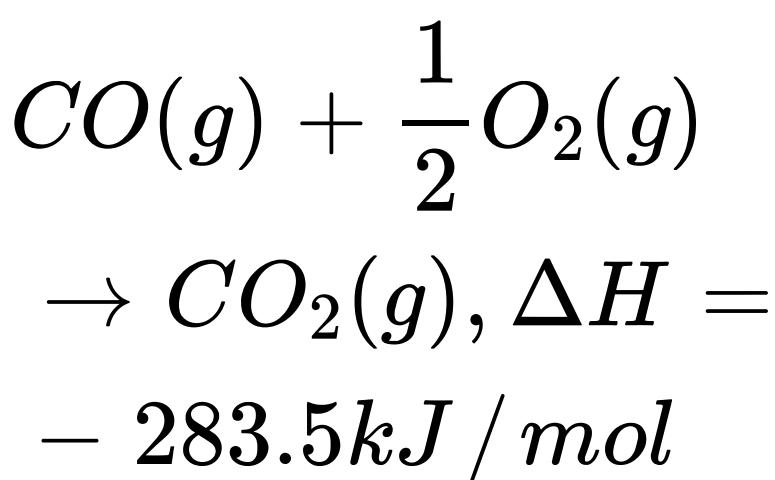
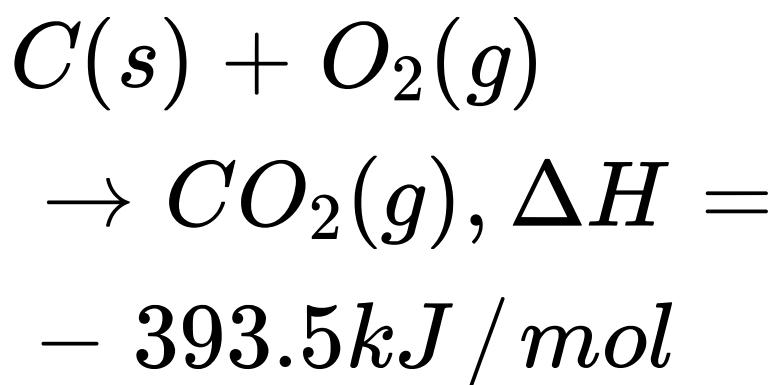
(D) 110.5

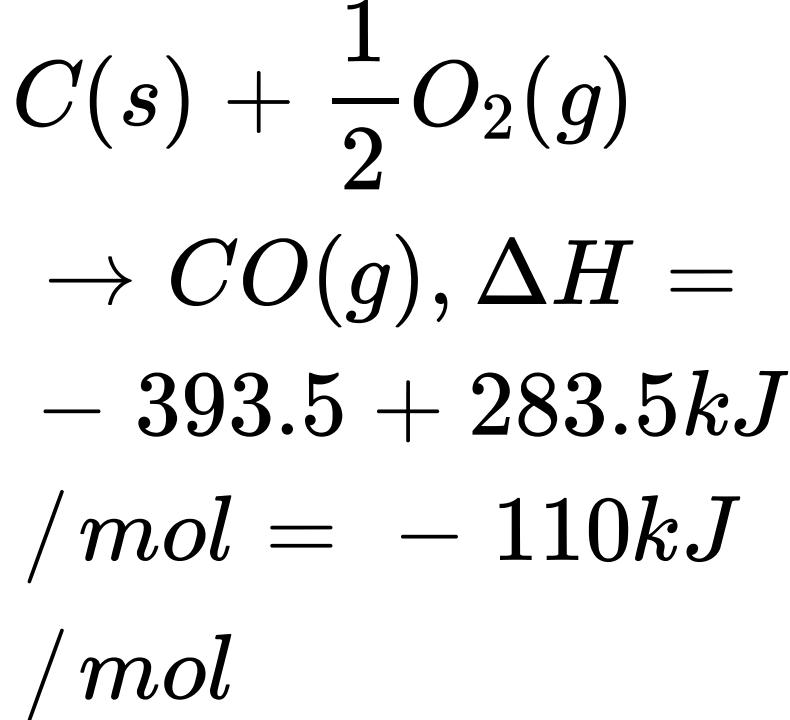
---

CORRECT ANSWER: C

---

SOLUTION:





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Q-39 - 52406166

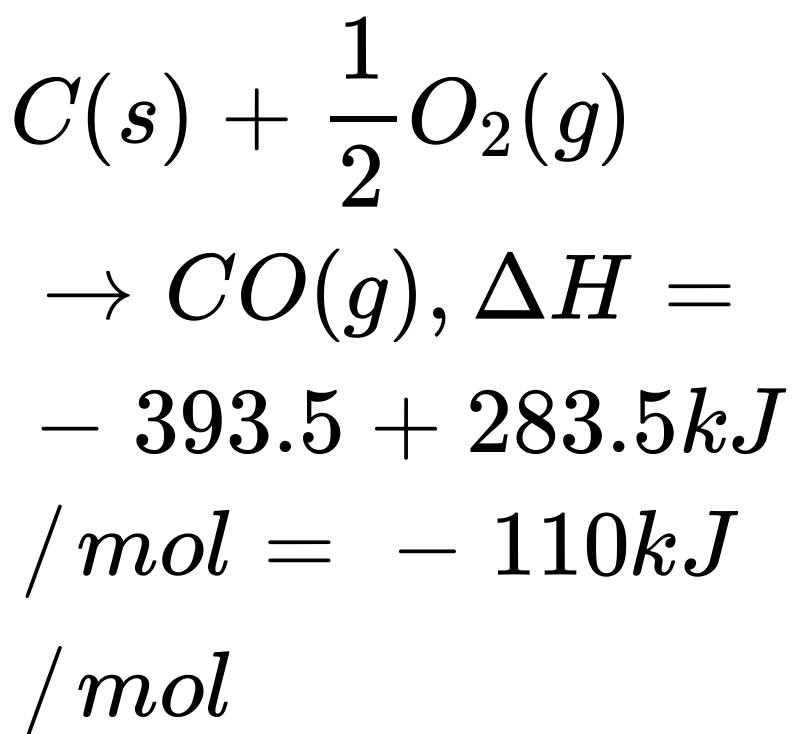
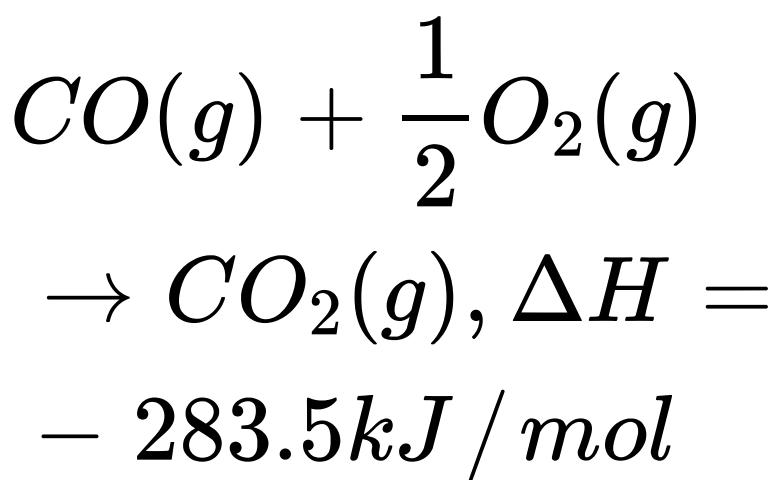
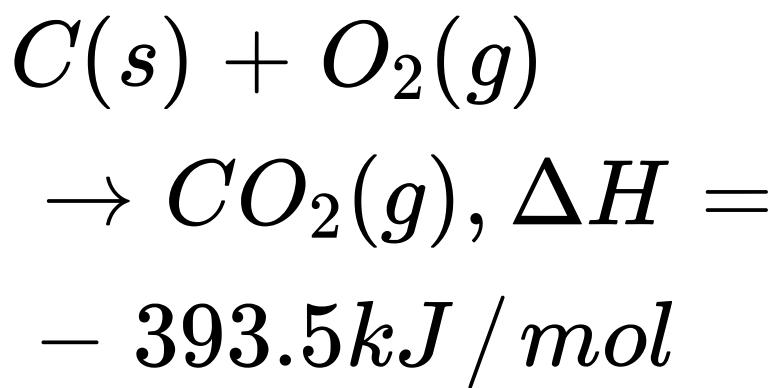
The heats of combustion of carbon and carbon monoxide are - 393.5 and  $-283.5 \text{ KJ mol}^{-1}$ , respectively. The heat of formation (in kJ) of carbon monoxide per mole is

- (A) 676.5
- (B) - 676.5
- (C) - 110.5
- (D) 110.5

## CORRECT ANSWER: C

---

SOLUTION:



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The combustion of benzene (1) gives  $CO_2(g)$  and  $H_2O(1)$ . Given that heat of combustion of benzene at constant volume is

$-3263.9\text{ kJ mol}^{-1}$  at  $25^\circ C$ , heat of combustion (in  $\text{kJ mol}^{-1}$ )

of benzene at constant pressure will be

$$(R = 8.314 \text{ J K}^{-1} \text{ mol}^{-1})$$

(A)  $-452.46$

(B)  $3260$

(C)  $-3267.6$

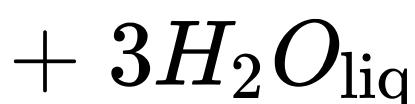
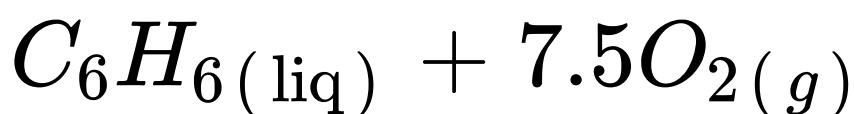
(D)  $4152.6$

---

CORRECT ANSWER: C

---

SOLUTION:



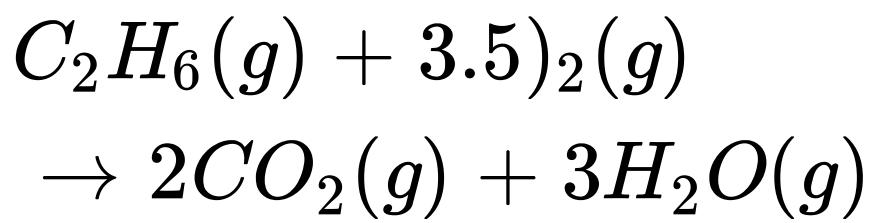
$$\Delta n_{(g)} = 6 - 7.5 = -15$$

$$\Delta H = \Delta E + \Delta n_{(g)} RT$$

$$\begin{aligned}\Delta H &= -3263.9 \text{ kJ} \\ &- \frac{1.5 \times 8.314 \times 298}{1000} \text{ kJ} \\ &= -3267.6 \text{ kJ}\end{aligned}$$

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Q-41 - 12225887



$$\Delta S_{vap}(H_2O, l) = x_1 cal K^{-1} \text{ (boiling point } = T_1)$$

$$\Delta H_f(H_2O, l) = x_2$$

$$\Delta H_f(CO_2) = x_3$$

$$\Delta H_f(C_2H_6) = x_4$$

Hence,  $\Delta H$  for the reaction is

(A)  $2x_3 + 3x_2 - x_4$

(B)

$$2x_3 + 3x_2 - x_4$$

$$+ 3x_1 T_1$$

(C)  $2x + 3x_2 - x_4 + 3x_1 T_1$

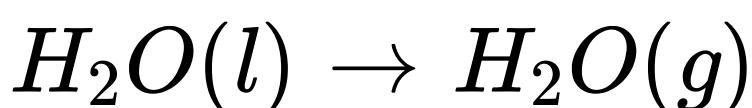
(D)  $x_1 T_1 + x_2 + x_3 - x_4$

---

CORRECT ANSWER: C

---

SOLUTION:



$$\Delta H_{vap} = \Delta S_{vap} T_{B.P.}$$

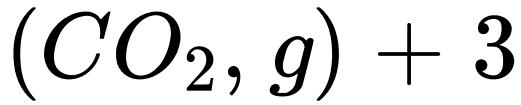
$$= x_1 T_1$$

$$\Delta H_f(H_2O, g) =$$

$$\Delta H_f(H_2O, l) + \Delta H_{vap}$$

$$= x_2 + x_1 T_1$$

$$\Delta H_{reaction} = 2\Delta H_f,$$



$$\Delta H_f(H_2O, g) -$$

$$\Delta H_f(C_2H_6, g)$$

$$= 2x_3 + 3(x_2 + x_1T_1)$$

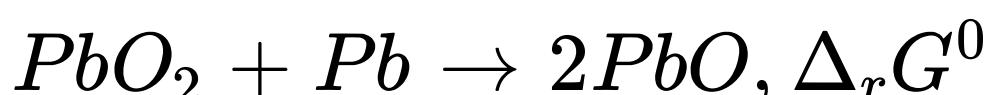
$$- x_4 = 2x_3 + 3x_2$$

$$+ 3x_1T_1 - x_4$$

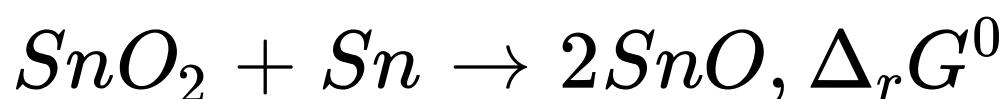
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Q-42 - 12225884

In view of the signs of  $\Delta_rG^0$  for the following reactions



$$< 0$$



$$> 0$$

Which oxidation state are more characteristic for lead and tin?

- (A) For lead +4 , for tin + 2

(B) For lead + 2 , for tin + 2

(C) For lead + 4 , for tin + 4

(D) For lead + 2 , for tin + 4

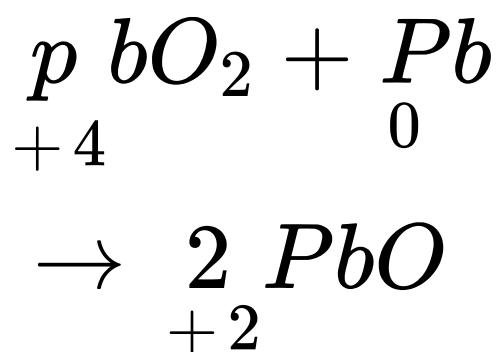
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CORRECT ANSWER: A

---

SOLUTION:

Oxidation state

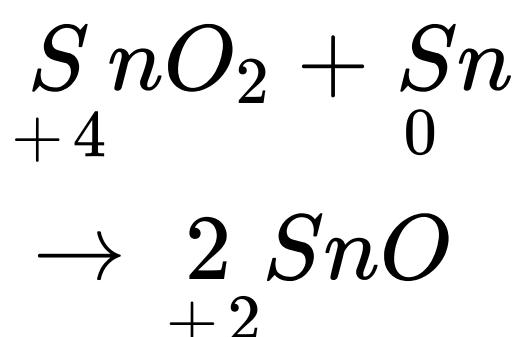


Since  $\Delta G < 0$  , i.e., it is negative. Therefore, the

reaction is spontaneous in the forward direction. This

suggest  $Pb^{2+}$  is more stable than  $Pb^{4+}$  .

Oxidation state

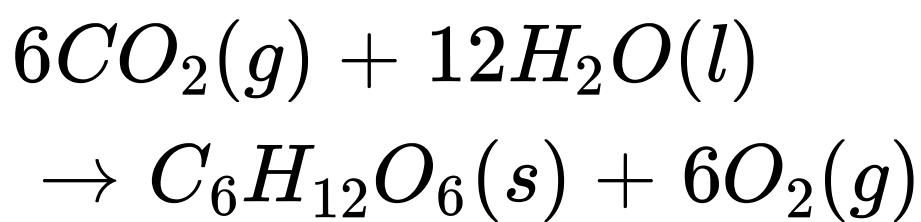


Since  $\Delta G > 0$ , i.e., it is positive, therefore, the reaction is non-spontaneous in the forward direction. But it will be spontaneous in the backward direction. This suggests that  $Sn^{2+}$ . These facts are also supported by the inert pair effect down the group.

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Q-43 - 12225882

For the reaction in the plant cells



$$+ 6H_2O(l)\Delta_r G^0 = 3000 KJ/mol$$



$$\Delta_r G^0 = -30 KJ/mol$$

Glucose is stored in the plant cell as starch  $(C_6H_{10}O_5)_n$ .

To produce  $162g$  of starch how many moles of  $ATP$  are minimum required?

(A)  $100mol$

(B)  $10mol$

(C)  $20mol$

(D)  $200mol$

---

CORRECT ANSWER: D

---

SOLUTION:

$$\text{Mol. Wt. of starch} = 162n$$

$$\therefore \text{Moles of } C_6H_{12}O_6 \text{ to be produced} = 1mol$$

$$\therefore \text{Moles of } ATP \text{ required} = \frac{3000}{30} = 100mol.$$

Which of the following statement (s) is/are correct?

Statement (*i*) : The entropy of isolated system with  $P - V$  work only, is always maximized at equilibrium

Statement (*ii*) : It is possible for the entropy of close system to decrease substantially in an irreversible process.

Statement (*iii*) : Entropy can be created but not destroyed.

Statement(iv) :  $\Delta S_{\text{system}}$  is zero for reversible process in an isolated system.

(A) Statement (*I, ii, iii*)

(B) Statement *ii, iv*

(C) Statement *I, ii, iv*

(D) All of these

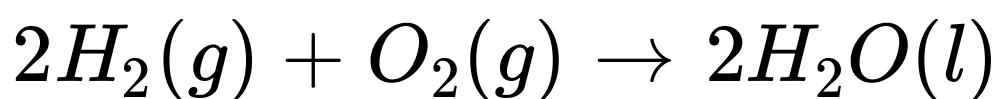
---

CORRECT ANSWER: D

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Q-45 - 12225909

The entropy change for the reaction given below is



is ... At 300K . Standard entropies of  $H_2(g)$ ,  $O_2(g)$  and  $H_2O(l)$  are 126.6, 201.20 and  $68.0\text{JK}^{-1}\text{mol}^{-1}$  respectively.

(A)  $-318.4\text{JK}^{-1}\text{mol}^{-1}$

(B)  $318.4 \times \text{JK}^{-1}\text{mol}^{-1}$

(C)  $31.84 \times \text{JK}^{-1}\text{mol}^{-1}$

(D) None of these

---

CORRECT ANSWER: A

---

SOLUTION:

$$\Delta S_{\text{Reaction}} = \Sigma S_{\text{product}}$$

$$- \Sigma S_{\text{reactant}}$$

$$= 2 \times S_{H_2O}$$

$$- [2 \times S_{H_2} + S_{O_2}]$$

$$= 2 \times 68$$

$$- [2 \times 126.6$$

$$+ 201.20] =$$

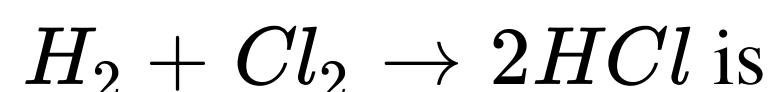
$$- 318.4 JK^{-1} mol^{-1}$$

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Q-46 - 12225913

If  $S$  for  $H_2$ ,  $Cl_2$  and  $HCl$  are  $0.13$ ,  $0.22$  and  $0.19 KJK^{-1} mol^{-1}$

respectively. The total change in standard entropy for the reaction



(A)  $0.5 KJK^{-1} mol^{-1}$

(B)  $40 JK^{-1} mol^{-1}$

(C)  $60JK^{-1}mol^{-1}$

(D)  $20JK^{-1}mol^{-1}$

---

CORRECT ANSWER: B

---

SOLUTION:

$$\begin{aligned}\Delta S &= S_P - S_R \\&= (2 \times 0.19) - 0.13 \\&\quad - 0.22 \\&= 0.03JK^{-1}mol^{-1} \\&= 30JK^{-1}mol^{-1}\end{aligned}$$

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Q-47 - 12225907

The standard entropies of  $N_2(g)$ ,  $H_2(g)$  and  $NH_3(g)$  are 191.5,  $192.6JK^{-1}mol^{-1}$ . The value of  $\Delta S^0$  during the formation of 1 mole of ammonia is

(A)  $-98.9 JK^{-1} mol^{-1}$

(B)  $-9.78 JK^{-1} mol^{(-1)}$

(C)  $+129.4 JK^{-1} mol^{-1}$

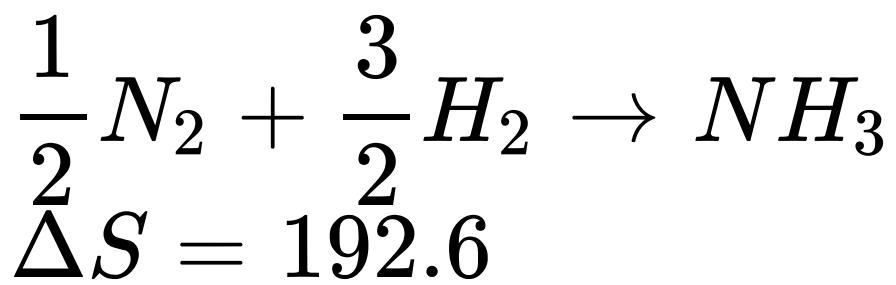
(D) Zero

---

CORRECT ANSWER: C

---

SOLUTION:



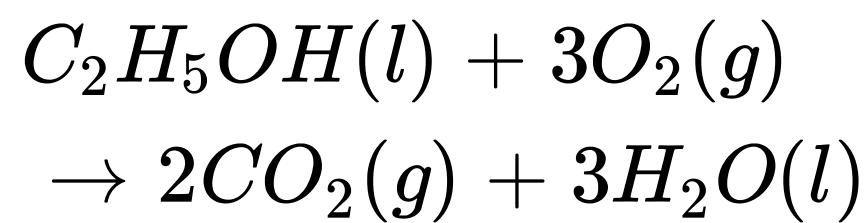
$$-\left(\frac{1}{2} \times 191.5 + \frac{3}{2} \times 130.5\right)$$

$$= 192.6 - (95.75 + 195.75)$$

$$= 192.6 - 291.5 =$$
$$-98.9 JK^{-1} mol^{-1}$$

Q-48 - 12225925

For the complete combustion of ethanol,



the amount of heat produced as measured in bomb calorimeter is

$1364.47\text{KJmol}^{-1}$  at  $25^\circ\text{C}$ . Assuming ideality, the enthalpy of

combustion,  $\Delta H_C$ , for the reaction will be

$$[R = 8.314\text{JK}^{-1}\text{mol}^{-1}]$$

(A)  $-1366.95\text{KJmol}^{-1}$

(B)  $-1361.95\text{KJmol}^{-1}$

(C)  $-1460.50\text{KJmol}^{-1}$

(D)  $-1350\text{KJmol}^{-1}$

---

CORRECT ANSWER: B

SOLUTION:



Amount of heat produced in bomb calorimeter,

$$\Delta U =$$

$$- 1364.47 KJmol^{-1}$$

Enthalpy of a combustion reaction is

$$\Delta H = \Delta U + \Delta_g RT$$

where  $\Delta U$  = internal energy

$\Delta n_g$  = moles of gas (products - reactants)

$\Delta R$  = Gas constant

$T$  = Temperature in  $K$

As per equation,

$$\Delta n_g = 2 - 3 = -1$$

$$T = 25C = 25$$

$$+ 273K$$

$$\Rightarrow T = 298K$$

Thus,

$$\Delta H = -1364.47$$

+

$$\left[ \frac{(-1) \times 8.314 \times 298}{1000} \right]$$

]

$$\Delta H = -1364.47$$

$$- 2.477 =$$

$$- 1366.947 KJmol^{-1}$$

Hence, the enthalpy of combustion,  $\Delta H_c$ , for the given reaction will be  $-1366.947 KJmol^{-1}$ .

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The amount of heat required to raise the temperature of 1 mole of diatomic gas by  $1^{\circ}\text{C}$  at constant pressure is  $60\text{cal}$ . The amount of heat which goes as internal energy of the gas is nearly.

(A)  $60\text{Cal}$

(B)  $30\text{Cal}$

(C)  $42.6\text{Cal}$

(D)  $49.8\text{Cal}$

---

CORRECT ANSWER: B

---

SOLUTION:

For a diatomic gas  $C_p = \frac{7}{2}R$  and  $C_v = \frac{5}{2}R$

Only  $\frac{5}{7} = 0.71$  of energy supplied increases the internal

energy of the gas. The rest is used to do work against

external pressure

$$0.71 \times 60 = 42.6\text{cal}$$

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Q-50 - 12225929

When  $20mL$  of a strong acid is added to  $20mL$  of an alkali, the temperature rises by  $5C$ . If  $200mL$  of each liquid is mixed, the temperature rise will be

(A)  $5^{\circ}C$

(B)  $50^{\circ}C$

(C)  $20^{\circ}C$

(D)  $0.5^{\circ}C$

---

CORRECT ANSWER: A

---

SOLUTION:

Heat evolved in the first case =  $40 \times 5 = 200\text{cal}$

Heat evolved in the second case = 10 times to the first,

that is  $10 \times 200 = 2000\text{cal}$

$$\therefore 2000 = 400 \times \Delta T \text{ or } \Delta T = 5C.$$

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Q-51 - 12225951

When 1 mole of gas is heated at constant volume. Temperature is raised from  $298$  to  $308K$ . Heat supplied to the gas is  $500J$ . Then which statement is correct?

(A)

$$q = -W = 500J,$$

$$\Delta U = 0$$

(B)

$$q = \Delta U = 500J, W = 0$$

(C)

$$q = -W = 500J,$$

$$\Delta U = 0$$

(D)

$$\Delta U = 0, q = W = \\ -500J$$

---

CORRECT ANSWER: C

SOLUTION:

We Know that,  $\Delta H = \Delta E + P\Delta V$

When,  $\Delta V = 0$ ,  $\therefore \Delta H = \Delta E$

From the first law of thermodynamics,

$$\Delta E = q - W$$

In the given problem,  $\Delta H = 500J$

$$\begin{aligned} -W &= -P\Delta V, \Delta V \\ &= 0 \end{aligned}$$

So,  $\Delta E = q = 500J$

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Two moles of ideal gas at  $27^\circ C$  temperature is expanded reversibly from 2 litre to 20 liter. Find entropy change ( $R = 2 \text{ cal/molK}$ ) .

(A) 92.1

(B) 0

(C) 4

(D) 9.2

---

CORRECT ANSWER: A

---

SOLUTION:

$\Delta S$  (Entropy change)

$$= 2.303nR \log_{10} \frac{V_2}{V_1}$$
$$= 2.303 \times 2$$

$$\times 2 \log_{10} \frac{20}{2}$$

$$= 2.303 \times 2 \times 2 \times 1$$

$$= 9.212 \text{ cal}$$

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Q-53 - 12225963

The densities of graphite and diamond at  $298K$  are  $2.25$  and  $3.31 \text{ g cm}^{-3}$ , respectively. If the standard free energy difference ( $\Delta G^0$ ) is equal to  $1895 \text{ J mol}^{-1}$ , the pressure at which graphite will be transformed into diamond at  $298K$  is

(A)  $9.92 \times 10^6 \text{ Pa}$

(B)  $9.92 \times 10^5 \text{ Pa}$

(C)  $9.92 \times 10^8 \text{ Pa}$

(D)  $9.92 \times 10^7 \text{ Pa}$

CORRECT ANSWER: B

SOLUTION:

$$\text{Volume of graphite} = \frac{\text{mass}}{\text{density}} = \frac{112}{2.25}$$

$$\text{Volume of diamond} = \frac{12}{3.31}$$

Change in volume,

$$\Delta V = \left( \frac{12}{3.31} - \frac{12}{2.25} \right) \times 10^{-3} L = -1.91 \times 10^{-3} L$$

$$\Delta G = \text{work done} = -p\Delta V$$

$$P = \frac{\Delta G}{\Delta V}$$
$$= \frac{1895 J mol^{-1}}{1.91 \times 10^{-3} \times 101.3}$$
$$= 9794$$

$$[ \because 1 \text{ atm} = 10^5 \times 1.013 \text{ Pa} ]$$

$$= 9.92 \times 10^8 \text{ Pa}$$

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Q-54 - 12225965

Standard enthalpy and standard entropy change for the oxidation of  $NH_3$  at  $298K$  are  $-382.64 KJmol^{-1}$  and  $145.6 Jmol^{-1}$  respectively. Standard free energy change for the same reaction at  $298K$  is

(A)  $-221.1 KJmol^{-1}$

(B)  $-339.3 KJmol^{-1}$

(C)  $-439.3 KJmol^{-1}$

(D)  $-523.2 KJmol^{-1}$

---

CORRECT ANSWER: A

---

SOLUTION:

$$\Delta H =$$
$$- 382.64 KJmol^{-1}$$

Given that,

$$\Delta S =$$
$$- 145.6 JK^{-1} mol^{-1}$$

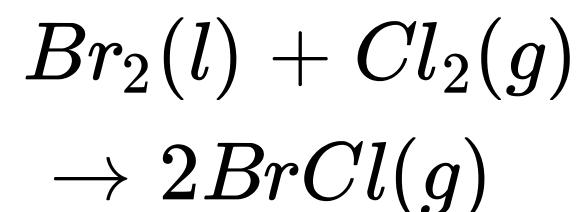
$$= - 145.6$$
$$\times 10^{-3} KJK^{-1}$$

or

$$\Delta G = - 382.64$$
$$- (298 \times 1 - 145.6$$
$$\times 10^3)$$
$$= - 339.3 KJmol^{-1} .$$

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The enthalpy and entropy change for the reaction



are  $30\text{ kJ mol}^{-1}$  and  $105\text{ J K}^{-1}\text{mol}^{-1}$  respectively. The

temperature at which the reaction will be in equilibrium is

(A) 450 K

(B) 300 K

(C) 285.7 K

(D) 273 K

---

CORRECT ANSWER: C

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Q-56 - 12225966

The work done during the expansion of a gas from a volume of

$4\text{dm}^3$  to  $6\text{dm}^3$  against a constant external pressure of  $3\text{atm}$  is

$$(1atm - L = 101.32J)$$

(A)  $-6J$

(B)  $-608J$

(C)  $+304J$

(D)  $-304J$

---

CORRECT ANSWER: B

---

SOLUTION:

Work done

$$(W) = - P_{ext.} (V_2) \\ - V_1)$$

$$= - 3 \times 6(6 - 4) =$$

$$- 6L - atm$$

$$= - 6 \times 101.32J( \\ \therefore 1Latm = 101.32J)$$

$$= - 607.92 \approx 608J.$$

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Q-57 - 12225967

The bond energies of  $H - - H$ ,  $Br - - Br$  and  $H - - Br$  are 433, , 192 and  $364 KJmol^{-1}$  respectively. The  $\Delta H$  for the reaction

$H_2(g) + Br_2(g) \rightarrow 2HBr(g)$  is

(A)  $261 KJ$

(B)  $+103 KJ$

(C)  $+261 KJ$

(D)  $-103 KJ$

---

CORRECT ANSWER: B

---

SOLUTION:

$$\Delta H = - 2 \times e_{H-Br}$$

$$+ e_{H-H} + e_{Br-Br}$$

$$= - 2 \times 364 + 344$$

$$+ 192 = - 103 KJ$$

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Q-58 - 11882429

The enthalpy of hydrogenation of cyclohexene is  $-119.5 \text{ kJ mol}^{-1}$

. If resonance energy of benzene is  $-150.4 \text{ kJ mol}^{-1}$ , its enthalpy of hydrogenation would be :

(A)  $-269.9 \text{ kJ mol}^{-1}$

(B)  $-358.5 \text{ kJ mol}^{-1}$

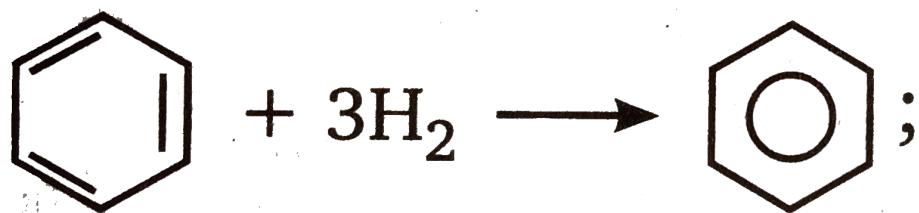
(C)  $-508.9 \text{ kJ mol}^{-1}$

(D)  $-208.1 \text{ kJ mol}^{-1}$

---

CORRECT ANSWER: D

## SOLUTION:



$$\begin{aligned}\Delta H &= -3 \\ &\times (119.5) \text{ kJ} = \\ &- 358.5 \text{ kJ}\end{aligned}$$

The Obs.  $\Delta H = \Delta H + \text{Energy needed to disturb resonance}$

$$\begin{aligned}&= -358.5 + 150.4 \\ &= -208.1 \text{ kJ mol}^{-1}\end{aligned}$$

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Q-59 - 23681941

From the following bond energies :-

$H - H$  bond energy

$$= 420 \text{ KJ mol}^{-1}$$

$C = -c$  bond energy

$$= 601 KJmol^{-1}$$

$C - C$  bond energy

$$= 340 KJmol^{-1}$$

$C - H$  bond energy

$$= 425 KJmol^{-1}$$

Enthalpy for the reaction :-



(A)  $-599 KJmol^{-1}$

(B)  $-580 KJmol^{-1}$

(C)  $-625 KJmol^{-1}$

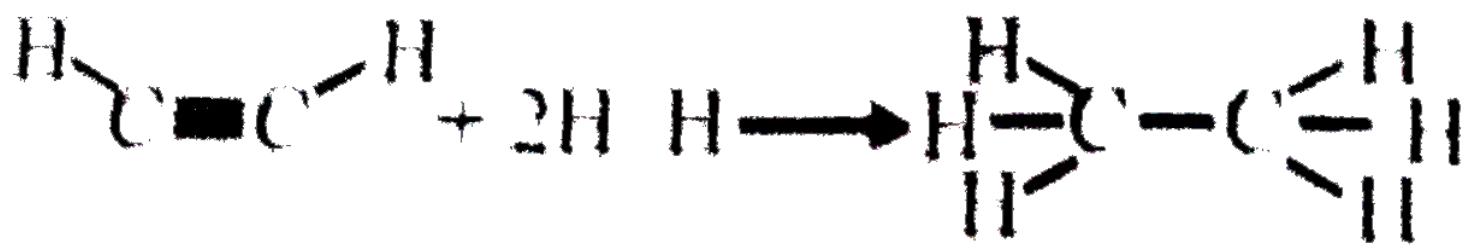
(D)  $-325 KJmol^{-1}$

---

CORRECT ANSWER: A

---

SOLUTION:



$$\Delta H_{rxn} = 2B$$

$$. E(C - H) + B$$

$$. E(C \equiv C) + 2B$$

$$. E(H - H) - 6B$$

$$. E(C - H) - B$$

$$. E(C - C)$$

$$\Rightarrow B. E(C \equiv C)$$

$$+ 2B. E(H - H)$$

$$- 4B. E(C - H) - B$$

$$. E(C - C)$$

$$\Rightarrow 601 + 2 \times 240 - 4$$

$$\times 425 - 340$$

$$\Delta H_{rxn} = -599 \text{ KJ/mol}$$

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For vaporization of water at 1 atmospheric pressure, the values of  $\Delta H$  and  $\Delta S$  are  $40.63 \text{ kJ mol}^{-1}$  and  $108.8 \text{ JK}^{-1}\text{mol}^{-1}$ , respectively. The temperature when Gibbs energy change ( $\Delta G$ ) for this transformation will be zero, is

(A) 273.4 K

(B) 393.4 K

(C) 373.4 K

(D) 293.4 K

---

CORRECT ANSWER: C

---

SOLUTION:

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

$$\Delta G = 0$$

$$\Delta H = T\Delta S,$$

$$T = \frac{40.63 \times 10^3}{108.8}$$

$$= 373.4K$$

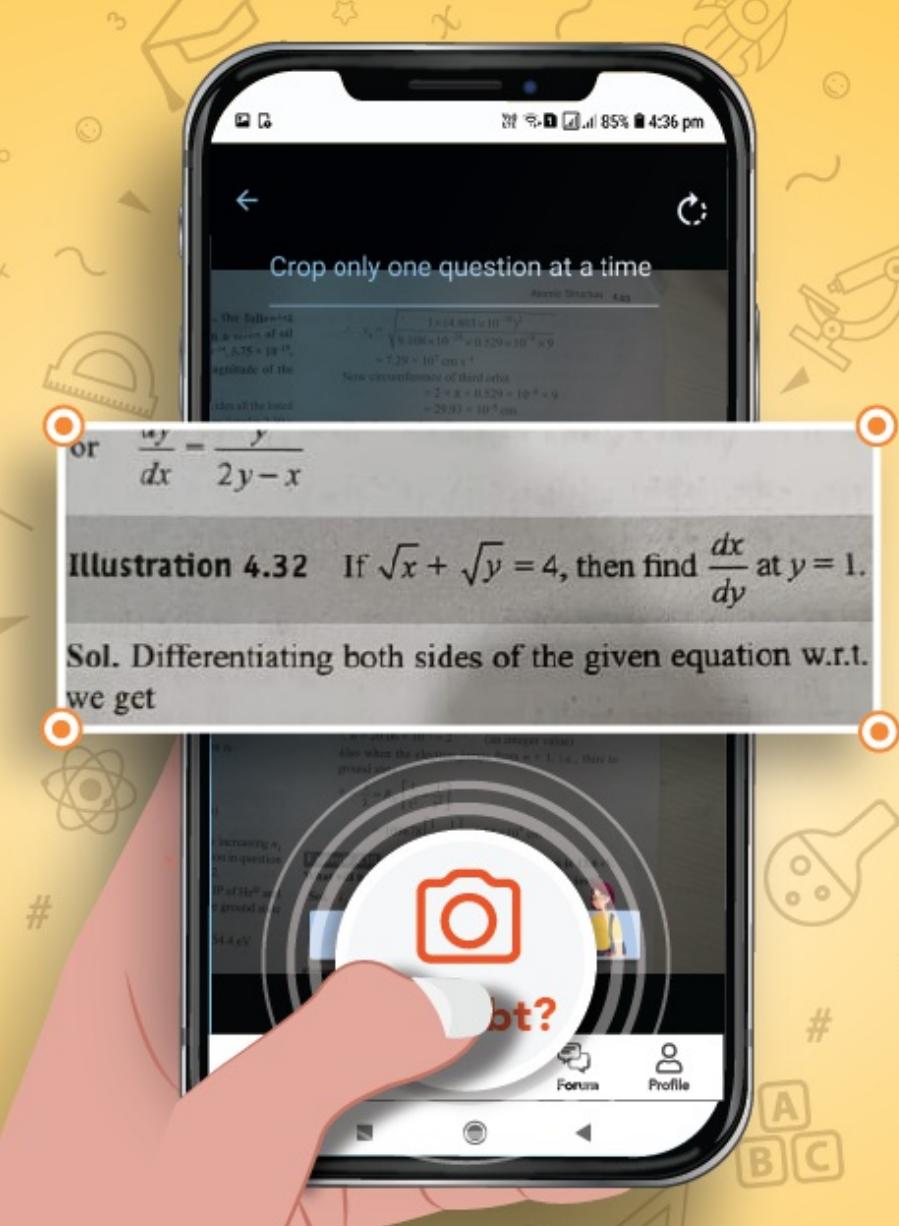
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