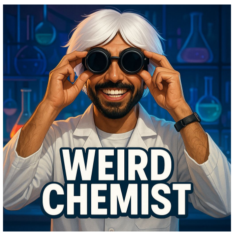

DPP–4 Solutions [Enthalpy of Combustion] (

The expert in anything was once a beginner.

Q1. The heat of combustion of a substance is –

> **Approach (Hinglish):** Combustion is exothermic for normal fuels $\Rightarrow \Delta H_c^\circ < 0$ (enthalpy decreases).

Final Answer: (B) Always negative

Q2. Correct expression of enthalpy of combustion

> **Approach (Hinglish):** Standard enthalpy of combustion is defined as the enthalpy change when **1 mole of a substance is completely burnt in oxygen** to form the most **stable oxidation products at 298 K and 1 atm.**
> For example:
> $$\mathrm{C \to CO_2(g)}, \quad \mathrm{H_2 \to H_2O(l)}, \quad \mathrm{S \to SO_2(g)}$$
> *Note:* "Complete oxidation" means oxidation till the thermodynamically stable product under standard conditions — not necessarily the **highest oxide**. Thus, although sulfur can form SO_3, the stable product at 298 K is SO_2(g), so that is used in the definition.

(A) forms CO (incomplete oxidation), (B) uses diamond (non-standard allotrope) and forms CO, (C) has wrong stoichiometry for CO_2.

Final Answer: (D) $\mathrm{S(s)_{rhombic} + O_2(g) \to SO_2(g)}$

Q3. Which equation corresponds to ΔH_c° at 298 K?

> **Approach (Hinglish):** At 298 K, water as **liquid**; definition usually per 1 mol of fuel.

Final Answer: (C) $\mathrm{C_2H_6(g) + \frac{7}{2}O_2 \to 2CO_2 + 3H_2O(l)}$

Q4. 0.4 g CH_4 gives 0.25 kCal. Find molar heat of combustion of CH_4.

> **Approach (Hinglish):** $n = \dfrac{0.4}{16} = 0.025$ mol. Heat per mol $= \dfrac{0.25}{0.025} = 10$ kCal mol^{-1} (released).

Final Answer: (B) 10 kCal

Q5. Benzene combustion: $\Delta H = -3264.40$ kJ mol^{-1}. Heat from 7.8 g?

> **Approach (Hinglish):** $n = \dfrac{7.8}{78} = 0.10 \text{ mol} \Rightarrow q = 0.10 \times 3264.40 = 326.44$ kJ (released).

Final Answer: (D) 326.4 kJ

Q6. Bomb calorimeter: X g ethanol releases Y J. Then —

Approach (Hinglish): Bomb calorimeter measures ΔE at constant V. For the *performed* burn on X g, energy released is Y J $\Rightarrow \Delta E = -Y$ J for that process. (Molar value would be $-\dfrac{MY}{X}$ with $M = 46$ g mol^{-1}.)

Final Answer: (B) $\Delta E_{\textbf{combustion}} = -Y$ J

Q7. 100 g glucose $\equiv 1560$ kJ. Energy per 1 g?

Approach (Hinglish): Direct proportion: $1560/100 = 15.6$ kJ g^{-1}.

Final Answer: (A) 15.6 kJ

Q8. 1 g anhydrous oxalic acid releases 2.835 kJ. ΔH_{comb} (molar)?

Approach (Hinglish): $M(H_2C_2O_4) \approx 90$ g mol^{-1}. Per mol: $2.835 \times 90 = 255.15$ kJ mol^{-1} (released).

Final Answer: (A) 255.15 kJ

Q9. Heat of formation of CO_2 is -94.0 kCal. Heat from burning 3 g C?

Approach (Hinglish): 12 g C $\to$ 94.0 kCal. For 3 g: $94.0 \times \dfrac{3}{12} = 23.5$ kCal.

Final Answer: (A) 23.5 kCal

Q10. ΔH_c°(C$\to$CO$_2$)$= -393.5$ kJ mol^{-1}. Heat on forming 35.2 g CO_2?

Approach (Hinglish): $n_{CO_2} = \dfrac{35.2}{44} = 0.80 \text{ mol} \Rightarrow q = 0.80 \times (-393.5) = -315$ kJ (approx).

Final Answer: (C) -315 kJ

Q11. Daily 2870 kCal; $\Delta H_c^{\circ\circ}$(sucrose)$= -1349$ kCal mol^{-1}. Sugar needed?

Approach (Hinglish): $n = \dfrac{2870}{1349} = 2.128$ mol; $M = 342$ g mol^{-1} $\Rightarrow$ mass ≈ 728 g.

Final Answer: (A) 728 g

Q12. Glucose: $\Delta H = -680$ kCal (per mol). Heat = 170 kCal. CO_2 made?

Approach (Hinglish): Fraction of a mole $= 170/680 = 1/4$. CO_2 formed $= 6 \times (1/4) = 1.5$ mol $\Rightarrow$ mass $= 1.5 \times 44 = 66$ g.

Final Answer: (C) 66 g

Q13. ΔH_f°: C_2H_2(g)$= +52$, $CO_2(g) = -394$, H_2O(l)$= -286$ kJ mol^{-1}. ΔH_c° of C_2H_2?

Approach (Hinglish): $C_2H_2 + \frac{5}{2}O_2 \to 2CO_2 + H_2O(l)$;

$$\Delta H_c^\circ = \big[2(-394) + (-286)\big] - [+52] = -1126 \text{ kJ mol}^{-1}.$$

Note: Given options show 1412 kJ; with the printed ΔH_f° values above, the rigorous result is -1126 kJ mol^{-1}. (Likely data/units mismatch in question.)

Final Answer: Closest conceptual choice: (B) *exothermic large magnitude*; computed value with given data: -1126 **kJ mol**$^{-1}$

Q14. $\Delta H_f^{\circ\circ}$**(CO)**$= -110$, $\Delta H_f^{\circ\circ}$**(CO**$_2$**)**$= -394$ **kJ mol**$^{-1}$**. Heat when 1 mol graphite burns?**

> **Approach (Hinglish):** $C + O_2 \rightarrow CO_2$; $\Delta H_{\text{rxn}} = \Delta H_f^{\circ\circ}(CO_2) - \Delta H_f^{\circ\circ}(C) = -394 - 0 = -394$ kJ.

Final Answer: (C) -394 **kJ**

Q15. ΔH_c°**: C(s)**$= -94$, **H**$_2(g) = -68$, **CH**$_4(g) = -213$ **kcal mol**$^{-1}$**.** ΔH **for C + 2H**$_2 \rightarrow$ **CH**$_4$**?**

> **Approach (Hinglish):** Use $\Delta H_{\text{rxn}} = \sum \Delta H_c^\circ(\text{react}) - \sum \Delta H_c^\circ(\text{prod}) = [(-94) + 2(-68)] - (-213) = -17$ kcal.

Final Answer: (A) -17 **kcal**

Q16. ΔH_c°**: C**$= -394$, **CO**$= -285$ **kJ mol**$^{-1}$**.** $\Delta H_f^{\circ\circ}$**(CO) ?**

> **Approach (Hinglish):** $C + O_2 \rightarrow CO_2$ minus $CO + \frac{1}{2}O_2 \rightarrow CO_2$ gives $C + \frac{1}{2}O_2 \rightarrow CO$.
> $\Rightarrow \Delta H_f^{\circ\circ}(CO) = -394 - (-285) = -109$ kJ.

Final Answer: (B) -109 **kJ mol**$^{-1}$

Q17. 2 g CH$_4$ **combustion** $\rightarrow$ **26575 Cal. Given** ΔH_f°**(CO**$_2$**)**$= -97000$ **Cal,** ΔH_f°**(H**$_2$**O)**$= -68000$ **Cal.** ΔH_f°**(CH**$_4$**)?**

> **Approach (Hinglish):** Per mol heat: $26575 \times \dfrac{16}{2} = 212600$ Cal $= \Delta H_c^\circ(CH_4)$.
> $\Delta H_c^\circ = [\Delta H_f^\circ(CO_2) + 2\Delta H_f^\circ(H_2O)] - \Delta H_f^\circ(CH_4)$
> $-212600 = (-97000 - 136000) - \Delta H_f^\circ(CH_4) \Rightarrow \Delta H_f^\circ(CH_4) = -20400$ Cal.

Final Answer: (C) -20400 **Cal**

Q18. ΔH_f°**(H**$_2$**O)**$= -285$, ΔH_f°**(CO**$_2$**)**$= -394$ **kJ mol**$^{-1}$**;** ΔH_c°**(CH**$_4$**)**$= -890$**.** ΔH_f°**(CH**$_4$**)?**

> **Approach (Hinglish):** $\Delta H_c^\circ = \{(-394) + 2(-285)\} - \Delta H_f^\circ(CH_4) \Rightarrow -890 = -964 - \Delta H_f^\circ(CH_4)$
> $\Rightarrow \Delta H_f^\circ(CH_4) = -74$ kJ mol^{-1}.

Final Answer: (D) -74 **kJ mol**$^{-1}$

Q19. ΔH_f°**(H**$_2$**O,l)**$= -68.3$ **kcal;** ΔH_c°**(C**$_2$**H**$_2$**)**$= -337.2$**;** ΔH_c°**(C**$_2$**H**$_4$**)**$= -363.7$**. For C**$_2$**H**$_2$ **+ H**$_2 \rightarrow$ **C**$_2$**H**$_4$

> **Approach (Hinglish):** Also $\Delta H_c^\circ(H_2) = \Delta H_f^\circ(H_2O) = -68.3$ (per mol).
> $\Delta H_{\text{rxn}} = \Delta H_c^\circ(C_2H_2) + \Delta H_c^\circ(H_2) - \Delta H_c^\circ(C_2H_4) = -337.2 - 68.3 + 363.7 = -41.8$ kcal.

Final Answer: (C) -41.8 **kCal**

Q20. If calorific value of C$_2$**H**$_6$ **is** x **kJ kg**$^{-1}$**, then** ΔH_c° **(per mol) is –**

> **Approach (Hinglish):** $M(C_2H_6) = 30$ g $= 0.030$ kg. Hence per mol heat $\approx x \times 0.030 = \dfrac{x}{30^{-1}} \approx \boxed{x/30}$ (kJ mol^{-1}).

Final Answer: (C) $x/30$

Q21. ΔH_c° (kJ/mol): Propane -2220, Butane -2878, Pentane -3537. Order of kJ/g?

Approach (Hinglish): Compare $|\Delta H_c^\circ|/M$: Propane $2220/44 \approx 50.5$; Butane $2878/58 \approx 49.6$; Pentane $3537/72 \approx 49.1$ kJ g^{-1}.

Final Answer: (A) Propane > Butane > Pentane

Q22. ΔH_c° (kCal/mol): CH_4 -212.8, C_2H_6 -373.0, C_2H_4 -337.0, C_2H_2 -310.5. Best fuel (per g)?

Approach (Hinglish): CH_4: $212.8/16 = 13.3$; C_2H_6: $373/30 = 12.4$; C_2H_4: $337/28 = 12.0$; C_2H_2: $310.5/26 = 11.9$ kCal g^{-1}.

Final Answer: (A) CH_4

Q23. ΔH_c° (kJ/mol): CH_4 -890, C_2H_4 -1411, C_2H_6 -1560. Lowest kJ/g?

Approach (Hinglish): CH_4: $890/16 = 55.6$; C_2H_6: $1560/30 = 52.0$; C_2H_4: $1411/28 = 50.4$ $\Rightarrow$ lowest $= C_2H_4$.

Final Answer: (C) C_2H_4

Q24. Ethylene: $\Delta H_c^\circ = 1411$ kJ mol^{-1}. If 6226 kJ evolved, O_2 volume at NTP?

Approach (Hinglish): $n(C_2H_4) = 6226/1411 = 4.414$ mol. Reaction: $C_2H_4 + 3O_2 \rightarrow 2CO_2 + 2H_2O$
$n(O_2) = 3 \times 4.414 = 13.242$ mol $\Rightarrow V = 13.242 \times 22.4 \approx 296.5$ L.

Final Answer: (B) 296.5 L

Q25. 112 L water gas (equal H_2, CO) at STP. Heat evolved?

Approach (Hinglish): 56 L each $\Rightarrow n = 56/22.4 = 2.5$ mol each.
$Q = 2.5(241.8) + 2.5(283) = 604.5 + 707.5 = 1312$ kJ (released).

Final Answer: (C) 1312 kJ

Q26. 5 L mix (CH_4, C_3H_8) burns using 16 L O_2 at STP. Heat released?

Approach (Hinglish): Let x L CH_4, $(5-x)$ L C_3H_8. O_2 need $= 2x + 5(5-x) = 25 - 3x = 16 \Rightarrow x = 3$ L.
Moles: CH_4: $3/22.4 = 0.1339$; C_3H_8: $2/22.4 = 0.0893$.
$Q = 0.1339(890) + 0.0893(2220) \approx 119.8 + 198.2 = 318$ kJ.

Final Answer: (C) 317 kJ (approx)

Q27. ΔH for $CH_4 + \frac{1}{2}O_2 \rightarrow CH_3OH$ is negative. If $\Delta H_c^\circ(CH_4) = x$, $\Delta H_c^\circ(CH_3OH) = y$, then —

Approach (Hinglish): $\Delta H_{\text{rxn}} = \Delta H_c^\circ(\text{react}) - \Delta H_c^\circ(\text{prod}) = x - y < 0 \Rightarrow x < y$.

Final Answer: (B) $x < y$

Q28. $\Delta H_c^\circ(CH_4) = 890$ kJ mol^{-1}. Heat from 2.4 L CH_4 at STP?

Approach (Hinglish): $n = 2.4/22.4 = 0.10714$ mol $\Rightarrow q = 0.10714 \times 890 \approx 95.3$ kJ.

Final Answer: (A) 95.3 kJ

Q29. ΔH for combustion of C(s) is -94.4 kCal. Heat of formation of CO(g)?

Approach (Hinglish): Need ΔH for $CO + \frac{1}{2}O_2 \rightarrow CO_2$ (or $\Delta H_f^\circ(CO_2)$) to apply Hess's law. Without it, value is indeterminate.

Final Answer: (D) More data required

Quick Reminders

(1) For Hess's manipulations with combustions: $\Delta H_{\text{rxn}} = \sum \Delta H_c^\circ(\text{Reactants}) - \sum \Delta H_c^\circ(\text{Products})$.

(2) Standard combustion at 298 K uses $\mathbf{H_2O(l)}$ as product unless specifically stated otherwise.

(3) Be consistent with units (kJ vs kCal) and phases (especially for water).