

Practice Questions on Thermochemistry (some less-than-optimal questions removed)

- A gas absorbs 0.0 J of heat and then performs 30.7 J of work. The change in internal energy of the gas is
 - 61.4 J
 - 30.7 J
 - 61.4 J
 - 30.7 J
 - none of these
- What is the kinetic energy of a 1.56-kg object moving at 94.0 km/hr?
 - 5.32×10^2 kJ
 - 6.89×10^3 kJ
 - 5.32×10^{-4} kJ
 - 1.06×10^3 kJ
 - 2.04×10^1 kJ
- Which of the following statements correctly describes the signs of q and w for the following exothermic process at $P = 1$ atm and $T = 370$ K?
$$\text{H}_2\text{O}(g) \rightarrow \text{H}_2\text{O}(l)$$
 - q and w are negative.
 - q is positive, w is negative.
 - q is negative, w is positive.
 - q and w are both positive.
 - q and w are both zero.
- For a particular process $q = 20$ kJ and $w = 15$ kJ. Which of the following statements is true?
 - Heat flows from the system to the surroundings.
 - The system does work on the surroundings.
 - $\Delta E = 35$ kJ
 - All of the above are true.
 - None of the above are true.
- One mole of an ideal gas is expanded from a volume of 1.00 liter to a volume of 8.93 liters against a constant external pressure of 1.00 atm. How much work (in joules) is performed on the surroundings? Ignore significant figures for this problem. ($T = 300$ K; $1 \text{ L}\cdot\text{atm} = 101.3$ J)
 - 402 J
 - 803 J
 - 2.41×10^3 J
 - 905 J
 - none of these

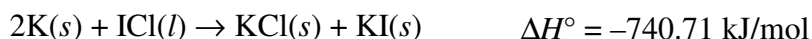
6. Calculate the work for the expansion of CO_2 from 1.0 to 4.7 liters against a pressure of 1.0 atm at constant temperature.
- A) 3.7 L·atm
 - B) 4.7 L·atm
 - C) 0 L·atm
 - D) -3.7 L·atm
 - E) -4.7 L·atm
7. Calculate the work associated with the compression of a gas from 121.0 L to 80.0 L at a constant pressure of 13.1 atm.
- A) -537 L atm
 - B) 537 L atm
 - C) 3.13 L atm
 - D) -3.13 L atm
 - E) 101 L atm
8. According to the first law of thermodynamics, the energy of the universe is constant. Does this mean that ΔE is always equal to zero?
- A) Yes, $\Delta E = 0$ at all times, which is why $q = -w$.
 - B) No, ΔE does not always equal zero, but this is only due to factors like friction and heat.
 - C) No, ΔE does not always equal zero because it refers to the system's internal energy, which is affected by heat and work.
 - D) No, ΔE never equals zero because work is always being done on the system or by the system.
 - E) No, ΔE never equals zero because energy is always flowing between the system and surroundings.
14. For the reaction $\text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{O}(g)$ at 298 K and 1.0 atm, ΔH is more positive than ΔE by 2.5 kJ. This quantity of energy can be considered to be
- A) the heat flow required to maintain a constant temperature
 - B) the work done in pushing back the atmosphere
 - C) the difference in the H-O bond energy in $\text{H}_2\text{O}(l)$ compared to $\text{H}_2\text{O}(g)$
 - D) the value of ΔH itself
 - E) none of these

15. Which one of the following statements is *false*?
- A) The change in internal energy, ΔE , for a process is equal to the amount of heat absorbed at constant volume, q_v .
 - B) The change in enthalpy, ΔH , for a process is equal to the amount of heat absorbed at constant pressure, q_p .
 - C) A bomb calorimeter measures ΔH directly.
 - D) If q_p for a process is negative, the process is exothermic.
 - E) The freezing of water is an example of an exothermic reaction.
16. $\text{C}_2\text{H}_5\text{OH}(l) + 3\text{O}_2(g) \rightarrow 2\text{CO}_2(g) + 3\text{H}_2\text{O}(l)$, $\Delta H = -1.37 \times 10^3 \text{ kJ}$
For the combustion of ethyl alcohol as described in the above equation, which of the following is true?
- I. The reaction is exothermic.
 - II. The enthalpy change would be different if gaseous water was produced.
 - III. The reaction is not an oxidation–reduction one.
 - IV. The products of the reaction occupy a larger volume than the reactants.
- A) I, II
 - B) I, II, III
 - C) I, III, IV
 - D) III, IV
 - E) only I
17. How much heat is required to raise the temperature of a 5.75-g sample of iron (specific heat = 0.450 J/g°C) from 25.0°C to 79.8°C?
- A) 2.54 J
 - B) 315 J
 - C) 700 J
 - D) 848 J
 - E) 142 J
18. A 45.9 g sample of a metal is heated to 95.2°C and then placed in a calorimeter containing 120.0 g of water ($c = 4.18 \text{ J/g}^\circ\text{C}$) at 21.6°C. The final temperature of the water is 24.5°C. Which metal was used?
- A) Aluminum ($c = 0.89 \text{ J/g}^\circ\text{C}$)
 - B) Iron ($c = 0.45 \text{ J/g}^\circ\text{C}$)
 - C) Copper ($c = 0.20 \text{ J/g}^\circ\text{C}$)
 - D) Lead ($c = 0.14 \text{ J/g}^\circ\text{C}$)
 - E) none of these

20. 30.0 mL of pure water at 282 K is mixed with 50.0 mL of pure water at 306 K. What is the final temperature of the mixture?
- A) 294 K
 - B) 297 K
 - C) 342 K
 - D) 588 K
 - E) 24 K
22. A 4.4-g sample of Colorado oil shale is burned in a bomb calorimeter, which causes the temperature of the calorimeter to increase by 5.0°C. The calorimeter contains 1.00 kg of water (heat capacity of H₂O = 4.184 J/g°C) and the heat capacity of the empty calorimeter is 0.10 kJ/°C. How much heat is released per gram of oil shale when it is burned?
- A) 21 kJ/g
 - B) 42 kJ/g
 - C) 0 kJ/g
 - D) 4.9 kJ/g
 - E) 0.21 kJ/g

23. If a student performs an endothermic reaction in a calorimeter, how does the calculated value of ΔH differ from the actual value if the heat exchanged with the calorimeter is not taken into account?
- A) ΔH_{calc} would be more negative because the calorimeter always absorbs heat from the reaction.
 - B) ΔH_{calc} would be less negative because the calorimeter would absorb heat from the reaction.
 - C) ΔH_{calc} would be more positive because the reaction absorbs heat from the calorimeter.
 - D) ΔH_{calc} would be less positive because the reaction absorbs heat from the calorimeter.
 - E) ΔH_{calc} would equal the actual value because the calorimeter does not absorb heat.

24. How much heat is liberated at constant pressure when 2.35 g of potassium metal reacts with 5.68 mL of liquid iodine monochloride ($d = 3.24 \text{ g/mL}$)?



- A) $2.22 \times 10^3 \text{ kJ}$
 - B) $8.40 \times 10^1 \text{ kJ}$
 - C) $1.28 \times 10^2 \text{ kJ}$
 - D) $2.23 \times 10^1 \text{ kJ}$
 - E) $7.41 \times 10^2 \text{ kJ}$
25. When 0.236 mol of a weak base (A^-) is reacted with excess HCl, 6.91 kJ of energy is released as heat. What is ΔH for this reaction per mole of A^- consumed?
- A) -34.2 kJ
 - B) -59.4 kJ
 - C) -29.3 kJ
 - D) 34.2 kJ
 - E) 29.3 kJ

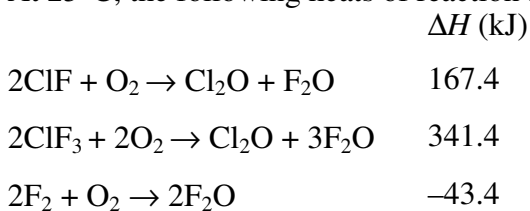
26. Consider the following numbered processes:

1. $\text{A} \rightarrow 2\text{B}$
2. $\text{B} \rightarrow \text{C} + \text{D}$
3. $\text{E} \rightarrow 2\text{D}$

ΔH for the process $\text{A} \rightarrow 2\text{C} + \text{E}$ is

- A) $\Delta H_1 + \Delta H_2 + \Delta H_3$
- B) $\Delta H_1 + \Delta H_2$
- C) $\Delta H_1 + \Delta H_2 - \Delta H_3$
- D) $\Delta H_1 + 2\Delta H_2 - \Delta H_3$
- E) $\Delta H_1 + 2\Delta H_2 + \Delta H_3$

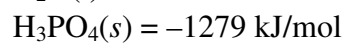
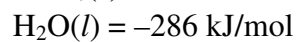
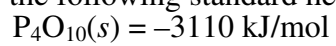
27. At 25°C, the following heats of reaction are known:



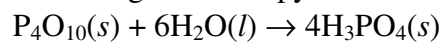
At the same temperature, calculate ΔH for the reaction: $\text{ClF} + \text{F}_2 \rightarrow \text{ClF}_3$

- A) -217.5 kJ
B) -130.2 kJ
C) +217.5 kJ
D) -108.7 kJ
E) none of these
28. The heat combustion of acetylene, $\text{C}_2\text{H}_2(g)$, at 25°C is -1299 kJ/mol. At this temperature, ΔH_f° values for $\text{CO}_2(g)$ and $\text{H}_2\text{O}(l)$ are -393 and -286 kJ/mol, respectively. Calculate ΔH_f° for acetylene.
- A) 2376 kJ/mol
B) 625 kJ/mol
C) 227 kJ/mol
D) -625 kJ/mol
E) -227 kJ/mol
29. For which of the following reaction(s) is the enthalpy change for the reaction *not* equal to ΔH_f° of the product?
- I. $2\text{H}(g) \rightarrow \text{H}_2(g)$
II. $\text{H}_2(g) + \text{O}_2(g) \rightarrow \text{H}_2\text{O}_2(l)$
III. $\text{H}_2\text{O}(l) + \text{O}(g) \rightarrow \text{H}_2\text{O}_2(l)$
- A) I
B) II
C) III
D) I and III
E) II and III

30. Consider the following standard heats of formation:



Calculate the change in enthalpy for the following process:



Answer Section

1. ANS: D
2. ANS: A
3. ANS: C
4. ANS: C
5. ANS: B
6. ANS: D
7. ANS: B
8. ANS: C

14. ANS: B
15. ANS: C
16. ANS: A
17. ANS: E
18. ANS: B

20. ANS: B

22. ANS: D

23. ANS: D

24. ANS: D

25. ANS: C

26. ANS: D

27. ANS: D

28. ANS: C

29. ANS: D

30. ANS:
-290 kJ

Solutions to Practice Questions for Ch. 6

1. $\Delta E = q + w = -30.7 \text{ J}$

\uparrow
0.0 J $\underbrace{-30.7}_{\text{negative}}$

because the system performs work (thus losing energy)

2. In addition to knowing the expression for kinetic energy, we also need to pay attention to units. SI units are:

mass: kg

length: m

time: s

energy: J (which is actually equivalent to $\text{kg} \cdot \frac{\text{m}^2}{\text{s}^2}$)

$$\text{K.E.} = \frac{1}{2} m v^2 = \frac{1}{2} (1.56 \text{ kg}) \left(94.0 \frac{\text{km}}{\text{h}} \times \frac{1000 \text{ m}}{1 \text{ km}} \times \frac{1 \text{ h}}{60 \text{ min}} \times \frac{1 \text{ min}}{60 \text{ s}} \right)^2 = 5.32 \times 10^2 \text{ J}$$

3. The question specifies that the process occurs at constant T and P, which means that V is the only variable responding to the changes happening due to the process (note that it is not a chemical reaction, but just a condensation of water vapor). It makes sense that it is exothermic, because condensing the vapor requires removal of heat. So, we are told that q is negative. Work, on the other hand, is positive because volume gets much smaller going from gas to liquid state. Remember, $w = -P\Delta V$.

4. $q = 20 \text{ kJ} > 0 \Rightarrow$ heat flows from the surroundings into the system
 $w = 15 \text{ kJ} > 0 \Rightarrow$ the surroundings does work on the system

$$\Delta E = q + w = 20 + 15 = 35 \text{ kJ}$$

$$5. w = -P\Delta V$$

Note that this equation doesn't have number of moles, n , or temperature in it. So the result doesn't depend on n or T that's given (redundant information).

$$w = -(1.00)(\overset{\substack{\uparrow \\ \text{final} \\ \text{volume}}}{8.93} - \overset{\substack{\uparrow \\ \text{initial} \\ \text{volume}}}{1.00}) = -7.93 \text{ L}\cdot\text{atm} \times \frac{101.3 \text{ J}}{1 \text{ L}\cdot\text{atm}} = -803 \text{ J}$$

$$\text{Work done on the surroundings} = -w = +803 \text{ J}$$

↑
defined as work done on the system
by the surroundings (thus the opposite
sign)

$$6. w = -P\Delta V = -(1.00)(4.7 - 1.0) = -3.7 \text{ L}\cdot\text{atm}$$

We use the regular formula for w , because we are not told anything specific about the direction of the work (on or by the system) so we assume its normal definition from the perspective of the system.

$$7. w = -P\Delta V = -(13.1)(80.0 - 121.0) = 537 \text{ L}\cdot\text{atm}$$

8. First law of thermodynamics is about the energy of the universe. The energy of a system is free to change, and ΔE does not have to be zero for a system (it can be, but it doesn't have to be).

13. Temperature and concentration are intensive since they don't change with the amount of substance (under the same conditions)

14.
$$\Delta H - \Delta E = 2.5 \text{ kJ}$$
$$\uparrow$$
$$\Delta E + \Delta(PV)$$

$$\Rightarrow \cancel{\Delta E} + \Delta(PV) - \cancel{\Delta E} = 2.5 \text{ kJ}$$
$$\Delta(PV) = 2.5 \text{ kJ}$$

@ const. P $\Delta(PV) = P\Delta V$ since P is not changing

$$P\Delta V = 2.5 \text{ kJ}$$

$$w = -P\Delta V$$

$$-w = P\Delta V$$

$$-w = 2.5 \text{ kJ}$$

$$w = -2.5 \text{ kJ}$$

So the system lost 2.5 kJ of energy due to work.

It performed 2.5 kJ of work by expanding against the pressure of the atmosphere when liquid H_2O turned into gaseous H_2O .

15. A) $\Delta E = q + w = q - \underbrace{P\Delta V}_{\text{at const. } V} = q_v$

B) $\Delta H = \Delta E + \underbrace{\Delta(PV)}_{P\Delta V \text{ @ const. } P} = q - P\Delta V + P\Delta V = q_p$

C) A bomb calorimeter operates at constant V, so $q_v = \Delta E$ not ΔH

D) $q_p < 0 \Rightarrow$ exothermic

E) To freeze water, q needs to be negative \Rightarrow exothermic

16. I. $\Delta H < 0 \Rightarrow$ exothermic

$$\text{II. } \Delta H = \Delta E + \Delta(PV)$$

PV term means that replacing liquid with gas (i.e. $\text{H}_2\text{O}(g)$ instead of $\text{H}_2\text{O}(l)$) with much larger volume would require more energy input to the system (for it to push back the atmosphere).

III. Combustion is an oxidation-reduction reaction. It's clear when we observe that elemental oxygen, O_2 , with \emptyset oxidation number ends up in compounds where its oxidation number cannot be zero.

IV. Volume is almost entirely determined by the number of moles of gas. The reaction starts with 3 moles of gas and ends up with 2 moles of gas. The products occupy a smaller volume.

17. $q = mc\Delta T = (5.75)(0.450)(79.8 - 25.0) = 142 \text{ J}$

By the way, we did not pay explicit attention to units because they were all consistent. But we do need to make sure they cancel to give a clean energy unit.

18. Heat lost by the hot metal is gained by the water.

$$-q_{\text{metal}} = q_{\text{water}}$$

$$-(45.9) c (24.5 - 95.2) = (120.0)(4.18)(24.5 - 21.6)$$

$$\Rightarrow c = 0.45 \text{ J/g}^\circ\text{C} \quad \text{the metal is iron}$$

20. heat lost by warm water = heat gained by cold water

$$-q_{\text{warm}} = q_{\text{cold}}$$

$$\left[\begin{array}{l} \text{remember:} \\ q = \text{heat gained} \\ -q = \text{heat lost} \end{array} \right]$$

We don't need to know the mass of waters explicitly, because they can be expressed in terms of volume and density, and densities cancel because we have the same substance on both sides of the equation. [Assuming density doesn't change much with T]

We also don't need to know the specific heat capacity of water, again because it shows up on both sides of the equation and cancels.

$$-\underbrace{V_{\text{warm}} \cancel{\rho_{\text{H}_2\text{O}}} / \cancel{c_{\text{H}_2\text{O}}}}_{\text{mass}_{\text{warm}}}(T_{\text{final}} - T_{\text{warm}}) = \underbrace{V_{\text{cold}} \cancel{\rho_{\text{H}_2\text{O}}} / \cancel{c_{\text{H}_2\text{O}}}}_{\text{mass}_{\text{cold}}}(T_{\text{final}} - T_{\text{cold}})$$

$$-(50.0)(T_{\text{final}} - 306) = (30.0)(T_{\text{final}} - 282)$$

$$\Rightarrow T_{\text{final}} = 297 \text{ K}$$

21. The "system" is composed of the reactants and the products the form. The solvent is the part of the surroundings immediately exchanging heat with the system. The solution is mainly reflecting what happens to the solvent. Since the solvent got colder, it lost energy, and the system gained energy. The chemical reaction is absorbing energy (choice C)

The way choice B is phrased, being very strictly technical, it would also have to be regarded correct (though it wasn't the intention of the writer of the question). Because:

$s \cdot m \cdot \Delta T$ is negative because ΔT is negative

"Energy released" being negative technically means energy was actually absorbed (it's like saying the amount of money lost was negative, meaning money was gained, or saying the amount of profit was negative, meaning money was lost)

And the magnitude of the heat absorbed is indeed given by the magnitude of $s \cdot m \cdot \Delta T$

22. The heat capacity of the calorimeter is that of the empty calorimeter plus the heat capacity of the water.

$$C = \underbrace{0.10 \times 10^3 \text{ J/}^\circ\text{C}}_{\text{heat capacity of empty calorimeter}} + \underbrace{(1.00 \times 10^3 \text{ g})(4.184 \text{ J/g}^\circ\text{C})}_{m_{\text{H}_2\text{O}} c_{\text{H}_2\text{O}} \text{ heat capacity of water}} = 4.28 \times 10^3 \text{ J/}^\circ\text{C}$$

$$q_{\text{released}} = C \Delta T = (4.28 \times 10^3) (5.0) = 2.14 \times 10^4 \text{ J}$$

$$q_{\text{released}} \text{ per gram} = \frac{q_{\text{released}}}{m_{\text{oil}}} = \frac{2.14 \times 10^4 \text{ J}}{4.4 \text{ g}} = 4.9 \times 10^3 \text{ J/g} = 4.9 \frac{\text{kJ}}{\text{g}}$$

23. Endothermic reaction absorbs heat q_{rxn} from the mixture in the calorimeter, lowering its temperature. The colder the mixture gets, the larger the magnitude of the ΔH_{rxn} .

$$\Delta H_{\text{rxn}} = q_{\text{rxn}} = -q_{\text{mixt.}} = -mc\Delta T$$

The heat transferred from the calorimeter would replace some of the heat lost to the reaction, making $q_{\text{mixt.}}$ appear smaller.

The magnitude of ΔT would be smaller (the mixture would not get as cold as the true ΔH_{rxn} would lead to)

ΔH_{calc} would be smaller.

In terms of the proper symbols and signs:

$$\Delta H_{\text{calc}} = -q_{\text{calc}} \leftarrow \text{apparent, calculated heat lost by the mixture}$$

$$q_{\text{calc}} = q_{\text{true}} + q_{\text{exch.}}$$

$\xrightarrow{(-)}$ mixture loses heat to the rxn. $(+)$ ← mixture gains heat from the calorimeter

$$\text{So, } |q_{\text{calc}}| < |q_{\text{true}}|$$

$$\text{Therefore } |\Delta H_{\text{calc}}| < |\Delta H_{\text{true}}|$$

24. First we need to determine the limiting reactant.

$$\text{moles of K} = 2.35 \text{ g K} \times \frac{1 \text{ mol K}}{39.10 \text{ g K}} = 0.0601 \text{ mol} \xrightarrow{\text{div. by its coefficient (2)}} 0.0300 \leftarrow \text{smaller} \Rightarrow \text{K is limiting}$$

$$\text{moles of ICl} = \underbrace{3.24 \text{ g}}_{\text{mass} = d \times V} \times 5.68 \text{ mL} \times \frac{1 \text{ mol ICl}}{162.36 \text{ g}} = 0.1133 \text{ mol} \xrightarrow{\text{div. by its coeff. (1)}} 0.1133$$

We treat ΔH° given with the reaction equation the same way as the substances involved. We can use reaction stoichiometry to find heat evolved or absorbed when a given amount of limiting reactant reacted.

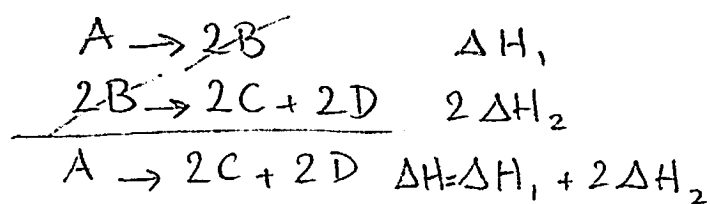
$$\text{Heat liberated} = -\Delta H \text{ for } 2.35 \text{ g K} = -2.35 \text{ g K} \times \frac{1 \text{ mol K}}{39.10 \text{ g K}} \times \frac{-740.71 \text{ kJ}}{2 \text{ mol K}} = 2.23 \times 10^1 \text{ kJ}$$

25. We don't need to construct the reaction equation. All we need to know is how much heat is released per mole of A^- . We also need to remember that "heat released" has the opposite sign of ΔH (a positive ΔH means heat is added to the system, while a positive "heat released" means heat is removed and the corresponding ΔH is negative).

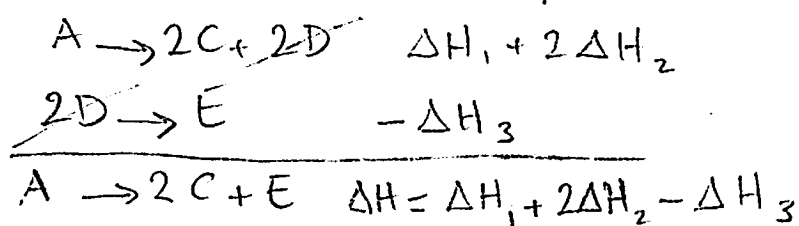
$$\Delta H = - \frac{6.91 \text{ kJ}}{0.236 \text{ mol } A^-} = -29.3 \text{ kJ/mol } A^-$$

26. In $A \rightarrow 2C + E$, A is a reactant (left side), so we don't need to reverse the direction of reaction 1. Also, in both the desired net reaction as well as in reaction 1 the coefficient of A is (implied) 1, so we don't need to multiply reaction 1 by any factor either.

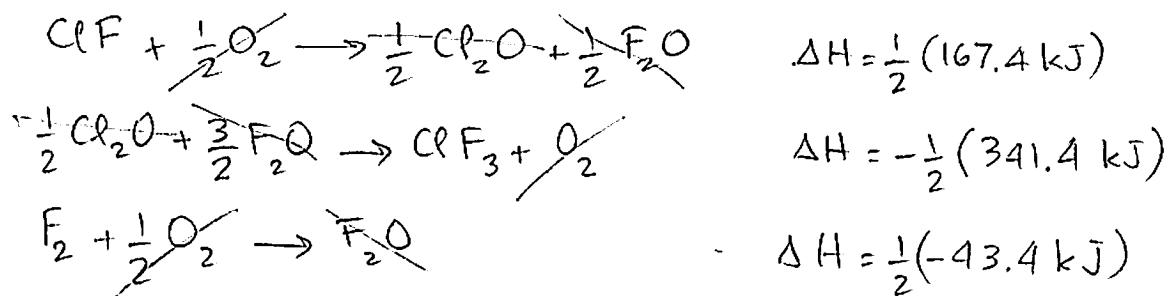
There is no B in the desired net reaction $A \rightarrow 2C + E$, so we need to cancel the $2B$ that is on the right side of reaction 1. We do that by multiplying reaction 2 by a factor of 2:



There is no D in the net desired reaction, and E is on the right side. We need to reverse the direction of reaction 3 and add it to the reaction equation obtained above

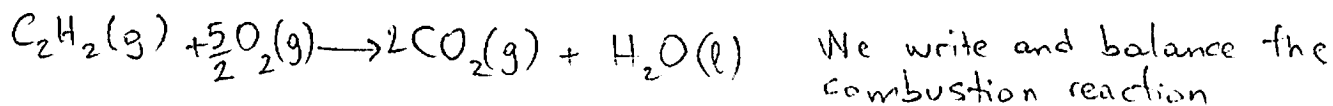


27. ClF and F_2 are on the left side of the desired net reaction, so we don't need to reverse the direction of reaction 1 or 3, but we do need to multiply each by $\frac{1}{2}$ because we want to get rid of the coefficient 2 in front of ClF and F_2 . We need to reverse the direction of reaction 2 and multiply by $\frac{1}{2}$ in order to have ClF_3 on the right side. Then we add them up and hope that the substances that are not in the net reaction cancel out.



Note that I used kJ instead of kJ/mol. I dislike using kJ/mol for reactions because "per mol" only makes sense if it's "per mole of reaction as written", which is more trouble than its worth as a concept. I prefer to simply write the reaction equation and the ΔH that corresponds to it, with the coefficients in the equation as written. No need to specify "per mol". You will encounter questions where the ΔH is carelessly given in kJ/mol, so that's why I am giving the above explanation rather than simply edit the question and getting rid of kJ/mol in it.

28.
$$\Delta H^\circ_{\text{rxn}} = \sum (\Delta H^\circ_f)_{\text{products}} - \sum (\Delta H^\circ_f)_{\text{reactants}}$$



$$\Delta H_{\text{rxn}} = \left[2(\Delta H^\circ_f)_{\text{CO}_2} + (\Delta H^\circ_f)_{\text{H}_2\text{O}} \right] - \left[(\Delta H^\circ_f)_{\text{C}_2\text{H}_2} + \frac{5}{2}(\Delta H^\circ_f)_{\text{O}_2} \right]$$

ΔH°_f values are always per mol, so we have to multiply by the coefficient in the rxn. eqn.

zero because O_2 is an element

$$-1299 = (2)(-393) + (-286) - \left[(\Delta H^\circ_f)_{\text{C}_2\text{H}_2} + 0 \right] \Rightarrow (\Delta H^\circ_f)_{\text{C}_2\text{H}_2} = 227 \text{ kJ/mol}$$

29. ΔH_f° is the ΔH° for the formation reaction of a compound. In a formation reaction, the only reactants are elements in their standard state, and the only product is one compound (and the coefficient of the product needs to be one)

I. $2\text{H}(g) \rightarrow \text{H}_2(g)$ not a formation reaction
Atomic H is not the standard state of H, and $\text{H}_2(g)$ does not have a formation reaction because it is already an element in its standard state.

II. $\text{H}_2(g) + \text{O}_2(g) \rightarrow \text{H}_2\text{O}_2(l)$ is a formation reaction with 1 mol of product, so its $\Delta H_{\text{rxn}}^\circ$ is equal to ΔH_f° for $\text{H}_2\text{O}_2(l)$

III. $\text{H}_2\text{O}(l) + \text{O}(g) \rightarrow \text{H}_2\text{O}_2(l)$ is not a formation reaction because one reactant is a compound, not an element, and the other reactant is an element that is not in its standard state (oxygen in its standard state is $\text{O}_2(g)$).

30.
$$\Delta H_{\text{rxn}}^\circ = \sum (\Delta H_f^\circ)_{\text{products}} - \sum (\Delta H_f^\circ)_{\text{reactants}}$$

$$\Delta H_{\text{rxn}}^\circ = (4)(-1279) - (-3110 + (6)(-286)) = -290 \text{ kJ}$$