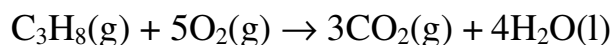


## Supplemental Practice Questions for Ch6

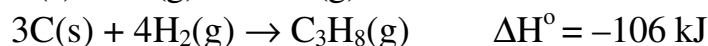
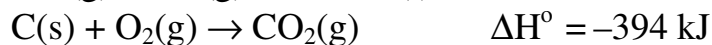
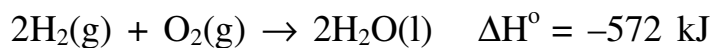
1. A piece of gold of mass 25.0 g at 88.0°C was placed in a calorimeter that contained 75.0 g of water at 25.00°C. If the temperature of the water rose to 25.70°C, what is the specific heat capacity of gold? The specific heat capacity of water is  $4.184 \text{ J}\cdot(\text{°C})^{-1}\cdot\text{g}^{-1}$ .

- A)  $0.10 \text{ J}\cdot(\text{°C})^{-1}\cdot\text{g}^{-1}$
- B)  $0.35 \text{ J}\cdot(\text{°C})^{-1}\cdot\text{g}^{-1}$
- C)  $0.14 \text{ J}\cdot(\text{°C})^{-1}\cdot\text{g}^{-1}$
- D)  $4.2 \text{ J}\cdot(\text{°C})^{-1}\cdot\text{g}^{-1}$
- E)  $5.2 \text{ J}\cdot(\text{°C})^{-1}\cdot\text{g}^{-1}$

2. Calculate the reaction enthalpy for the combustion of propane,  $\text{C}_3\text{H}_8(\text{g})$ ,

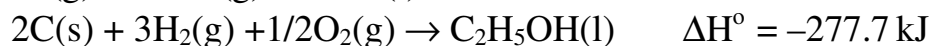
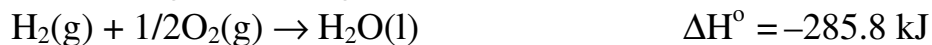


from the data



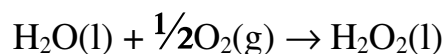
- A) -106 kJ
- B) -1072 kJ
- C) -2326 kJ
- D) -2220 kJ
- E) -2432 kJ

3. Calculate the reaction enthalpy for the combustion of 1 mole of ethanol,  $\text{C}_2\text{H}_5\text{OH}(\text{l})$ , from the data



- A) -1366.7 kJ
- B) -348.1 kJ
- C) -70.4 kJ
- D) -1644.4 kJ
- E) -1922.1 kJ

4. If the standard enthalpy of combustion of octane,  $\text{C}_8\text{H}_{18}(\text{l})$ , is  $-5471 \text{ kJ} \cdot \text{mol}^{-1}$ , calculate the enthalpy change when 1.00 kg of octane are burned.
- A)  $-2736 \text{ kJ}$
  - B)  $-23.9 \times 10^3 \text{ kJ}$
  - C)  $-5471 \text{ kJ}$
  - D)  $-47.9 \times 10^3 \text{ kJ}$
  - E)  $-95.8 \times 10^3 \text{ kJ}$
5. All of the following reactions or processes are endothermic except
- A) combustion.
  - B) vaporization.
  - C) sublimation.
  - D) melting.
  - E)  $\text{I}_2(\text{s}) \rightarrow \text{I}_2(\text{g})$
6. The standard enthalpy of formation of  $\text{H}_2\text{O}_2(\text{l})$  is  $-187.8 \text{ kJ} \cdot \text{mol}^{-1}$  and of  $\text{H}_2\text{O}(\text{l})$  is  $-285.8 \text{ kJ} \cdot \text{mol}^{-1}$ . Calculate the reaction enthalpy for



- A)  $-187.8 \text{ kJ}$
  - B)  $-473.6 \text{ kJ}$
  - C)  $-98.0 \text{ kJ}$
  - D)  $+187.8 \text{ kJ}$
  - E)  $+98.0 \text{ kJ}$
7. The reaction enthalpy for
- $$2\text{SnO}_2(\text{s}) \rightarrow 2\text{SnO}(\text{s}) + \text{O}_2(\text{g})$$
- is  $+589.8 \text{ kJ}$ . The standard enthalpy of formation of  $\text{SnO}(\text{s})$  is  $-285.8 \text{ kJ} \cdot \text{mol}^{-1}$ . What is the standard enthalpy of formation of  $\text{SnO}_2(\text{s})$ ?
- A)  $-1161.4 \text{ kJ} \cdot \text{mol}^{-1}$
  - B)  $-580.7 \text{ kJ} \cdot \text{mol}^{-1}$
  - C)  $-304.0 \text{ kJ} \cdot \text{mol}^{-1}$
  - D)  $+580.7 \text{ kJ} \cdot \text{mol}^{-1}$
  - E)  $-608.0 \text{ kJ} \cdot \text{mol}^{-1}$

Solutions to Supplemental Practice Questions for Ch. 6

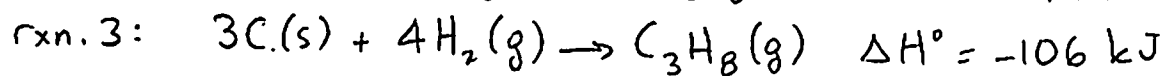
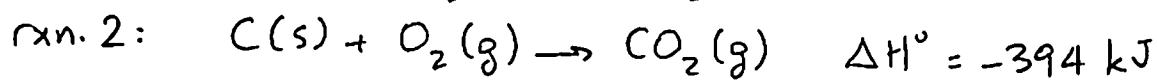
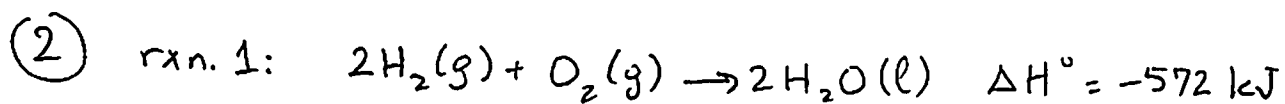
$$\textcircled{1} \quad q_{\text{Au}} = -q_{\text{H}_2\text{O}}$$

$T_f = 25.70^\circ\text{C}$  is the same for both gold and water

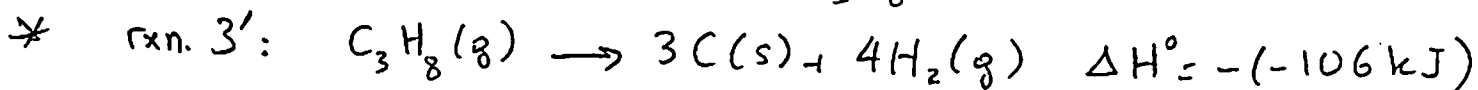
$$m_{\text{Au}} C_{\text{Au}} [T_f - (T_i)_{\text{Au}}] = -m_{\text{H}_2\text{O}} C_{\text{H}_2\text{O}} [T_f - (T_i)_{\text{H}_2\text{O}}]$$

$$(25.0) C_{\text{Au}} (25.70 - 88.0) = -(75.0)(4.184)(25.70 - 25.00)$$

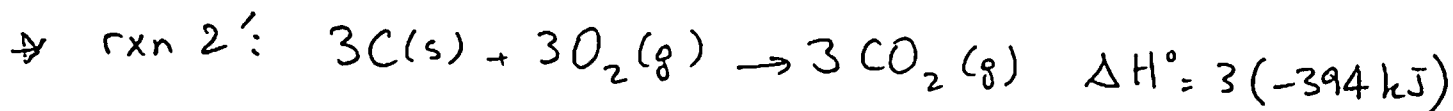
$$C_{\text{Au}} = 0.141 \text{ J/g}\cdot^\circ\text{C}$$



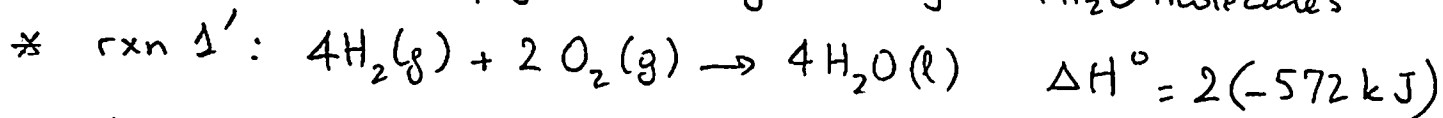
We start with a reaction that involves  $\text{C}_3\text{H}_8$  (rxn. 3), and reverse its direction so that  $\text{C}_3\text{H}_8$  is a reactant.



Rxn. 2 is the only reaction that gives  $\text{CO}_2$ , so we keep its direction, but multiply it by 3 because the desired net reaction gives 3  $\text{CO}_2$  molecules.



Rxn. 1 is the only one involving  $\text{H}_2\text{O}$ , which we need as a product. We need to multiply rxn. 1 by 2 to get 4  $\text{H}_2\text{O}$  molecules



Adding rxns 3', 2', and 1', we obtain the desired reaction and its  $\Delta H^\circ$ .

$$\Delta H^\circ_{\text{net}} = -(-106) + 3(-394) + 2(-572) = -2220 \text{ kJ}$$

② Alternatively, we recognize that the reactions given are the formation reactions of the compounds involved, except the first one has to be divided by 2 in order to have 1 mol of  $\text{H}_2\text{O}(\text{l})$  as product. So we have:

$$\Delta H_{\text{f}, \text{H}_2\text{O}(\text{l})}^{\circ} = -572 \text{ kJ}/2 \text{ mol} = -286 \text{ kJ/mol}$$

$$\Delta H_{\text{f}, \text{CO}_2(\text{g})}^{\circ} = -394 \text{ kJ}/1 \text{ mol} = -394 \text{ kJ/mol}$$

$$\Delta H_{\text{f}, \text{C}_3\text{H}_8(\text{g})}^{\circ} = -106 \text{ kJ}/1 \text{ mol} = -106 \text{ kJ/mol}$$

Then we can simply use

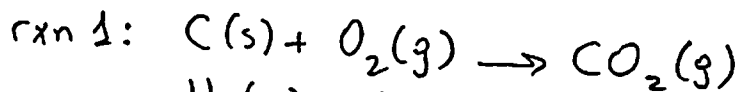
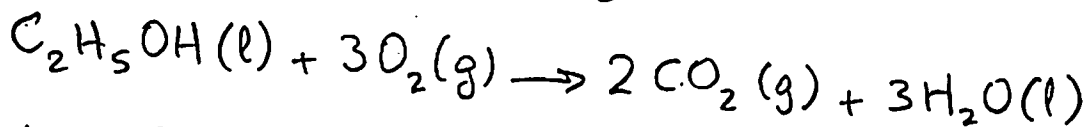
$$\Delta H_{\text{rxn}}^{\circ} = \sum_{\text{products}} n_i \Delta H_{\text{f}, i}^{\circ} - \sum_{\text{reactants}} n_i \Delta H_{\text{f}, i}^{\circ}$$

$$\Delta H_{\text{rxn}}^{\circ} = [(3)(-394) + (4)(-286)] - [(1)(-106) + (5)(0)] = -2220 \text{ kJ}$$

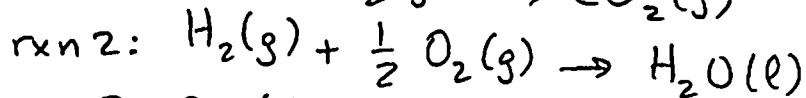
The advantage here is that the equation above takes care of multiplying, reversing, and adding of chemical reactions automatically and implicitly.

Of course we can only use it if the reactions given are basically formation reactions, as they were here.

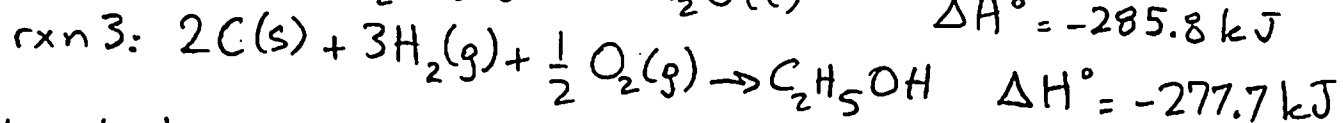
(3) Combustion reaction of  $C_2H_5OH$  is



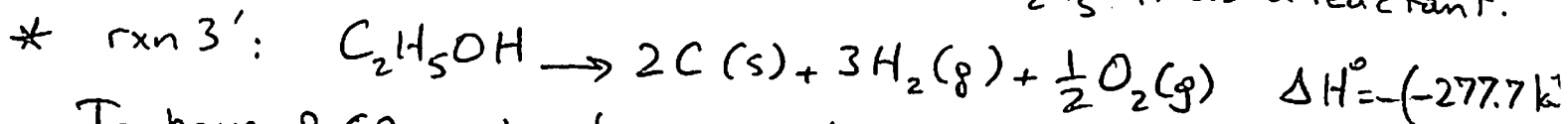
$$\Delta H^\circ = -393.5 \text{ kJ}$$



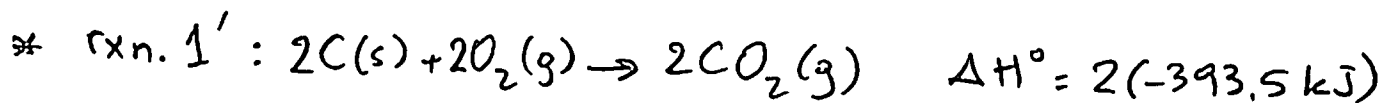
$$\Delta H^\circ = -285.8 \text{ kJ}$$



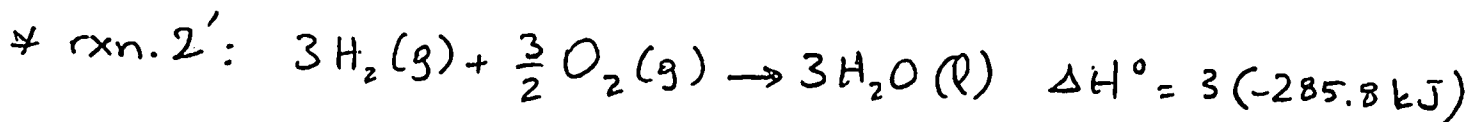
We start with reversing rxn. 3 to have  $C_2H_5OH$  as a reactant.



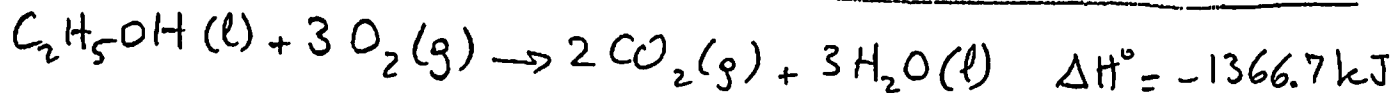
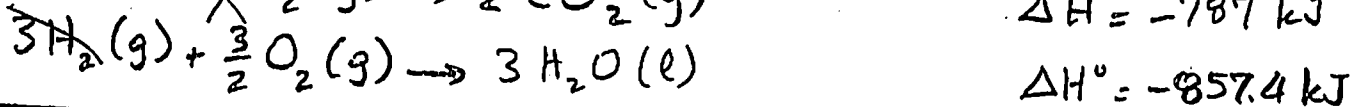
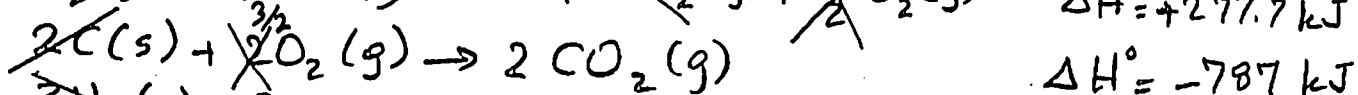
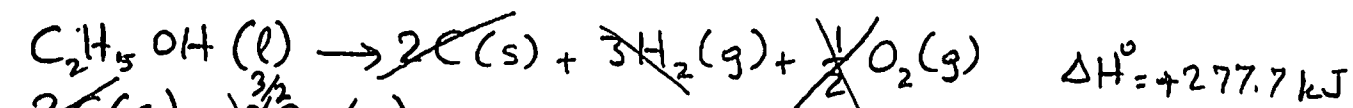
To have 2  $CO_2$  molecules as product, we need to multiply rxn. 1 by 2



To have 3  $H_2O$  molecules as product, we need to multiply rxn. 2 by 3



Adding rxns 3', 1', and 2',



Alternatively, we recognize that the reactions given are formation reactions, and we can use  $\Delta H^\circ_{\text{rxn}} = \sum n_i \Delta H^\circ_{f,i} \text{ (products)} - \sum n_i \Delta H^\circ_{f,i} \text{ (reactants)}$ , which applies Hess's Law implicitly.

$$\Delta H^\circ_{\text{rxn}} = [(2)(-393.5) + (3)(-285.8)] - [(1)(-277.7) + (3)(0)] = -1366.7 \text{ kJ}$$

- ④ Combustion is synonymous with burning. We know  $\Delta H$  for 1 mol of octane burning. We calculate it for 1.00 kg.

$$\Delta H = 1.00 \text{ kg} \times \frac{1000 \text{ g}}{1.00 \text{ kg}} \times \frac{1 \text{ mol}}{114.23 \text{ g}} \times \frac{-5471 \text{ kJ}}{1 \text{ mol}} = -4.79 \times 10^4 \text{ kJ (equal to } -47.9 \times 10^3 \text{ kJ)}$$

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- ⑤ All the phase transitions in the choices are endothermic (going from low-energy phases to high-energy phases: liq.  $\rightarrow$  gas, solid  $\rightarrow$  liq., solid  $\rightarrow$  gas).  
Combustion is exothermic.
- 

- ⑥ We apply  $\Delta H_{\text{rxn}} = \sum_{\text{products}} n_i \Delta H_{f,i} - \sum_{\text{reactants}} n_i \Delta H_{f,i}$

$$\Delta H_{\text{rxn}} = (1) \Delta H_{f, \text{H}_2\text{O}_2} - \left[ (1) \Delta H_{f, \text{H}_2\text{O}} + \left( \frac{1}{2} \right) \underbrace{\Delta H_{f, \text{O}_2}}_{\text{zero}} \right] = (-187.8) - (-285.8) = 98 \text{ kJ}$$

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- ⑦ We use  $\Delta H_{\text{rxn}} = \sum_{\text{products}} n_i \Delta H_{f,i} - \sum_{\text{reactants}} n_i \Delta H_{f,i}$

This time we know  $\Delta H_{\text{rxn}}$ , and use the equation to solve for an unknown  $\Delta H_{f,i}$ :

$$\Delta H_{\text{rxn}} = \left[ (2)(-285.8) + (1)(\emptyset) \right] - \left[ (2) \Delta H_{f, \text{SnO}_2} \right] = 589.8 \text{ kJ}$$

$\uparrow$   
 $\Delta H_{f, \text{O}_2}$

$$\Delta H_{f, \text{SnO}_2} = \frac{589.8 - (2)(-285.8)}{(-2)} = -580.7 \text{ kJ/mol}$$