ANSWER TO "TRY YOURSELF" PROBLEM FROM STUDY SECTION 5.7

Try Yourself 5.7 a

Given the following data:

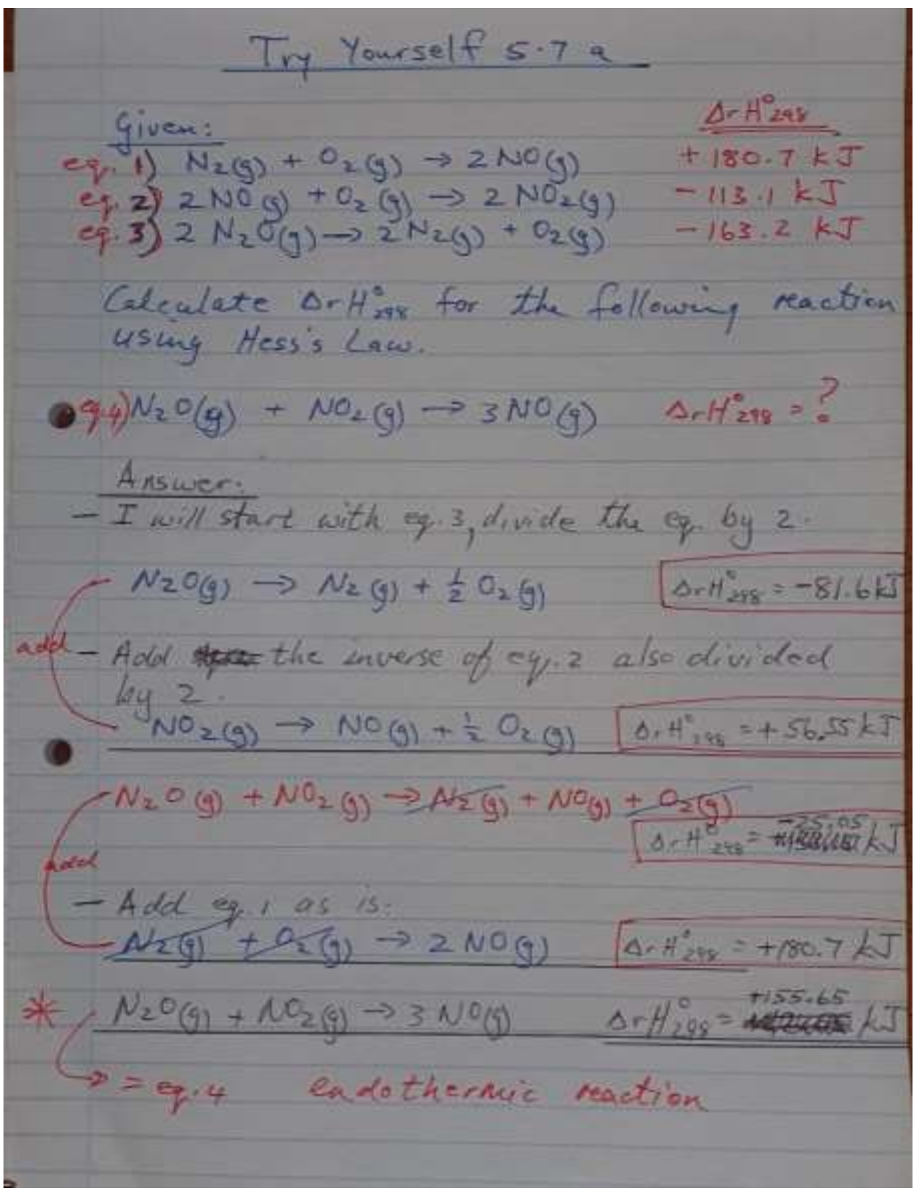
1. $N_2(g) + O_2(g) \rightarrow 2NO(g)$ $\Delta_r H^0_{298} = +180.7$ kJ

2. $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$ $\Delta_r H^0_{298} = -113.1$ kJ

3. $2N_2O(g) \rightarrow 2N_2(g) + O_2(g)$ $\Delta_r H^0_{298} = -163.2$ kJ

Calculate $\Delta_r H^0_{298}$ for the following reaction by using Hess's law and manipulating the given reactions above:

4. $N_2O(g) + NO_2(g) \rightarrow 3NO(g)$ $\mathbf{\Delta_r H^0_{298} = ?}$



Try Yourself 5·7 a

Given: $\Delta_r H^\circ_{298}$

eq. 1) $N_2(g) + O_2(g) \rightarrow 2NO(g)$ $+180.7$ kJ

eq. 2) $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$ -113.1 kJ

eq. 3) $2N_2O(g) \rightarrow 2N_2(g) + O_2(g)$ -163.2 kJ

Calculate $\Delta_r H^\circ_{298}$ for the following reaction using Hess's Law.

eq. 4) $N_2O(g) + NO_2(g) \rightarrow 3NO(g)$ $\Delta_r H^\circ_{298} = ?$

Answer:

– I will start with eq. 3, divide the eq. by 2.

$N_2O(g) \rightarrow N_2(g) + \frac{1}{2} O_2(g)$ $\boxed{\Delta_r H^\circ_{298} = -81.6 \text{ kJ}}$

add – Add the inverse of eq. 2 also divided by 2.

$NO_2(g) \rightarrow NO(g) + \frac{1}{2} O_2(g)$ $\boxed{\Delta_r H^\circ_{298} = +56.55 \text{ kJ}}$

$N_2O(g) + NO_2(g) \rightarrow \cancel{N_2(g)} + NO(g) + \cancel{O_2(g)}$ $\boxed{\Delta_r H^\circ_{298} = -25.05 \text{ kJ}}$

add – Add eq. 1 as is:

$\cancel{N_2(g)} + \cancel{O_2(g)} \rightarrow 2NO(g)$ $\boxed{\Delta_r H^\circ_{298} = +180.7 \text{ kJ}}$

* $N_2O(g) + NO_2(g) \rightarrow 3NO(g)$ $\Delta_r H^\circ_{298} = +155.65$ kJ

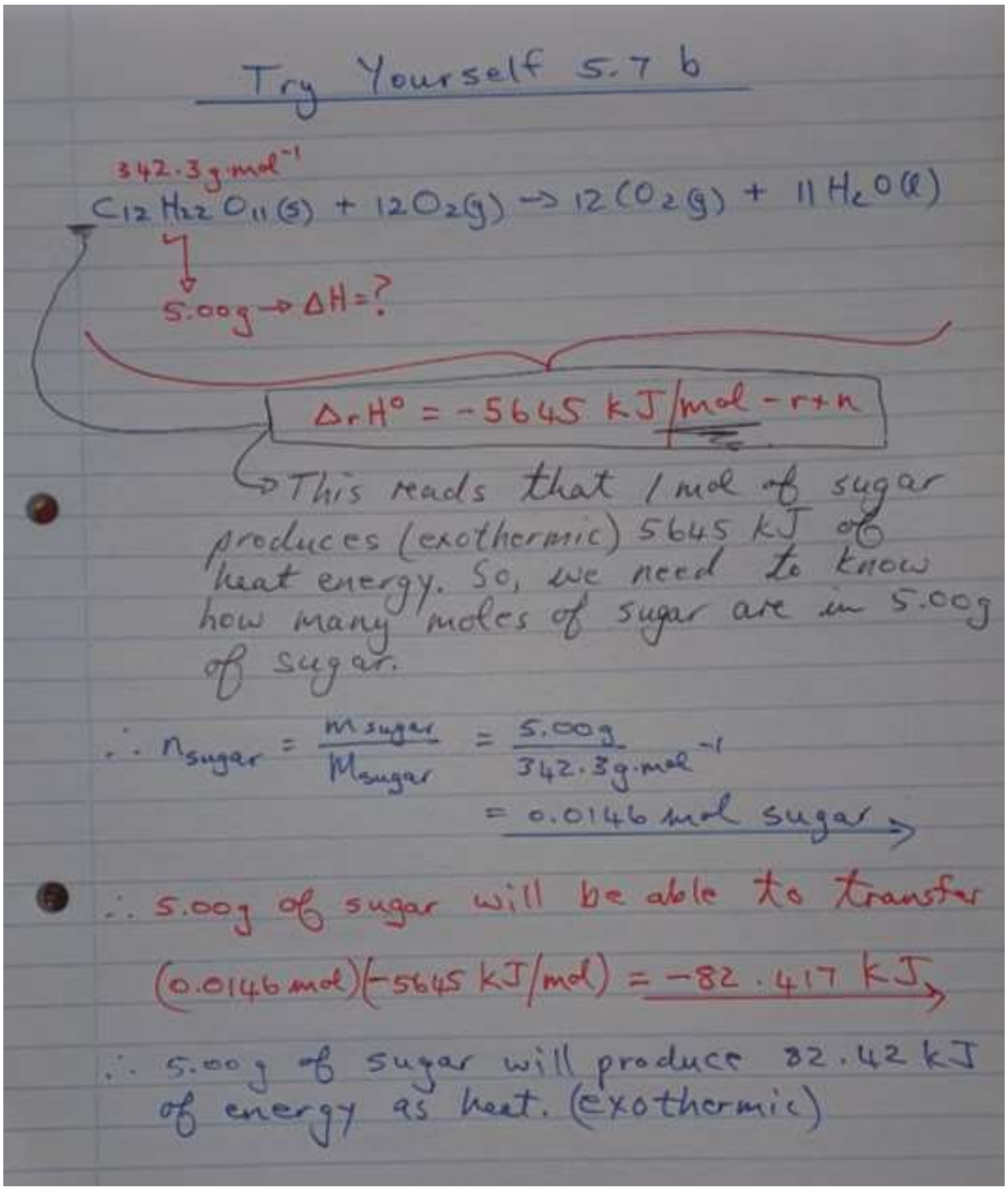
= eq. 4   endothermic reaction

Try Yourself 5.7 b

Sucrose (sugar, $C_{12}H_{22}O_{11}$ = 342.3 g.mol^{-1}) can be oxidized to CO_2 and H_2O and the enthalpy change for the reaction can be measured under conditions of constant pressure.

$$\mathbf{C_{12}H_{22}O_{11}(s) + 12O_2(g) \rightarrow 12CO_2(g) + 11H_2O(liq.) \qquad \Delta_rH^0 = -5645\ kJ/mol\text{-}rxn}$$

Calculate the energy that is transferred as heat by burning 5.00 g of sugar.



Try Yourself 5.7 b

$$\overset{342.3\,g\cdot mol^{-1}}{C_{12}H_{22}O_{11}(s)} + 12O_2(g) \rightarrow 12CO_2(g) + 11H_2O(l)$$

5.00g → ΔH = ?

$$\boxed{\Delta_rH^o = -5645\ kJ/\underline{mol}-rxn}$$

This reads that 1 mol of sugar produces (exothermic) 5645 kJ of heat energy. So, we need to know how many moles of sugar are in 5.00g of sugar.

$$\therefore n_{sugar} = \frac{m_{sugar}}{M_{sugar}} = \frac{5.00g}{342.3g\cdot mol^{-1}} = 0.0146\ mol\ sugar$$

∴ 5.00g of sugar will be able to transfer

$$(0.0146\ mol)(-5645\ kJ/mol) = -82.417\ kJ$$

∴ 5.00g of sugar will produce 82.42 kJ of energy as heat. (exothermic)

Try Yourself 5.7 c

Iso-octane (2,2,4-trimethylpentane = 114.18 g.mol^{-1}), one of the many hydrocarbons that make up gasoline, burns in air to give water and carbon dioxide.

$2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(l)$ Δ_rH^o = -10922 kJ-rxn

Calculate the enthalpy change if you burn 1.00 L of iso-octane (density = 0.69 g/mL).

Try Yourself 5.7 c

114.18 g.mol

$2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(l)$

1.00 L $\rightarrow \Delta H = ?$

1.00 L = 690 g

$\boxed{\Delta_r H^\circ = -10922 \text{ kJ-rxn}}$

This reads that 2 mol of C_8H_{18} produces (exothermic) 10922 kJ of heat energy. Therefore; 1 mol of C_8H_{18} will produce 5461 kJ of heat energy.

∴ 1.00 L of C_8H_{18} = 690 g C_8H_{18}

$$n_{C_8H_{18}} = \frac{m_{C_8H_{18}}}{M_{C_8H_{18}}} = \frac{690\text{ g}}{114.18\text{ g}\cdot\text{mol}^{-1}} = 6.04\text{ mol}$$

∴ 1.00 L of C_8H_{18} will produce:

$$6.04\text{ mol}\left(-5461\text{ kJ/mol}\right) = -32984.44\text{ kJ}$$