

# Enthalpy of Reaction ( $\Delta H_{rxn}$ )



$$\Delta H_{rx} = (2 \text{ mol})(\Delta H_f \text{CO}_2) - (2 \text{ mol})(\Delta H_f \text{CO}) - (1 \text{ mol})(\Delta H_f \text{O}_2)$$

$$\Delta H_{rx} = (2 \text{ mol})(-393.5 \text{ kJ/mol}) - (2 \text{ mol})(-110.5 \text{ kJ/mol}) - (1 \text{ mol})(0 \text{ kJ/mol})$$

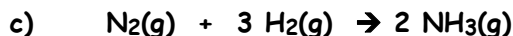
$$\underline{\underline{\Delta H_{rx} = -566.0 \text{ kJ}}}$$



$$\Delta H_{rx} = (1 \text{ mol})(\Delta H_f \text{CuSO}_4) + (1 \text{ mol})(\Delta H_f \text{SO}_2) + (2 \text{ mol})(\Delta H_f \text{H}_2\text{O(l)}) - (1 \text{ mol})(\Delta H_f \text{Cu}) - (2 \text{ mol})(\Delta H_f \text{H}_2\text{SO}_4)$$

$$\Delta H_{rx} = (1 \text{ mol})(-771.4 \text{ kJ/mol}) + (1 \text{ mol})(-296.8 \text{ kJ/mol}) + (2 \text{ mol})(-285.8 \text{ kJ/mol}) - (1 \text{ mol})(0 \text{ kJ/mol}) - (2 \text{ mol})(-814.0 \text{ kJ/mol})$$

$$\underline{\underline{\Delta H_{rx} = -11.8 \text{ kJ}}}$$



$$\Delta H_{rx} = (2 \text{ mol})(\Delta H_f \text{NH}_3) - (1 \text{ mol})(\Delta H_f \text{N}_2) - (3 \text{ mol})(\Delta H_f \text{H}_2)$$

$$\Delta H_{rx} = (2 \text{ mol})(-46.1 \text{ kJ/mol}) - (1 \text{ mol})(0 \text{ kJ/mol}) - (3 \text{ mol})(0 \text{ kJ/mol})$$

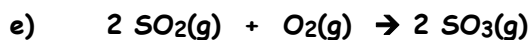
$$\underline{\underline{\Delta H_{rx} = -92.2 \text{ kJ}}}$$



$$\Delta H_{rx} = (2 \text{ mol})(\Delta H_f \text{Fe}) + (3 \text{ mol})(\Delta H_f \text{CO}_2) - (1 \text{ mol})(\Delta H_f \text{Fe}_2\text{O}_3) - (3 \text{ mol})(\Delta H_f \text{CO})$$

$$\Delta H_{rx} = (2 \text{ mol})(0 \text{ kJ/mol}) + (3 \text{ mol})(-393.5 \text{ kJ/mol}) - (1 \text{ mol})(-824.2 \text{ kJ/mol}) - (3 \text{ mol})(-110.5 \text{ kJ/mol})$$

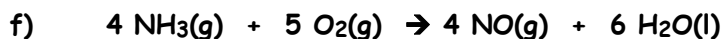
$$\underline{\underline{\Delta H_{rx} = -24.8 \text{ kJ}}}$$



$$\Delta H_{rx} = (2 \text{ mol})(\Delta H_f \text{SO}_3) - (2 \text{ mol})(\Delta H_f \text{SO}_2) - (1 \text{ mol})(\Delta H_f \text{O}_2)$$

$$\Delta H_{rx} = (2 \text{ mol})(-395.7 \text{ kJ/mol}) - (2 \text{ mol})(-296.8 \text{ kJ/mol}) - (1 \text{ mol})(0 \text{ kJ/mol})$$

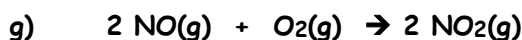
$$\underline{\underline{\Delta H_{rx} = -197.8 \text{ kJ}}}$$



$$\Delta H_{rx} = (4 \text{ mol})(\Delta H_f \text{NO}) + (6 \text{ mol})(\Delta H_f \text{H}_2\text{O(l)}) - (4 \text{ mol})(\Delta H_f \text{NH}_3) - (5 \text{ mol})(\Delta H_f \text{O}_2)$$

$$\Delta H_{rx} = (4 \text{ mol})(90.2 \text{ kJ/mol}) + (6 \text{ mol})(-285.8 \text{ kJ/mol}) - (4 \text{ mol})(-46.1 \text{ kJ/mol}) - (5 \text{ mol})(0 \text{ kJ/mol})$$

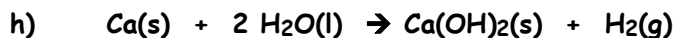
$$\underline{\underline{\Delta H_{rx} = -1169.6 \text{ kJ}}}$$



$$\Delta H_{\text{rx}} = (2 \text{ mol})(\Delta H_{\text{f}} \text{NO}_2) - (2 \text{ mol})(\Delta H_{\text{f}} \text{NO}) - (1 \text{ mol})(\Delta H_{\text{f}} \text{O}_2)$$

$$\Delta H_{\text{rx}} = (2 \text{ mol})(33.2 \text{ kJ/mol}) - (2 \text{ mol})(90.2 \text{ kJ/mol}) - (1 \text{ mol})(0 \text{ kJ/mol})$$

$$\underline{\underline{\Delta H_{\text{rx}} = -114.0 \text{ kJ}}}$$

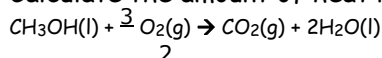


$$\Delta H_{\text{rx}} = (1 \text{ mol})(\Delta H_{\text{f}} \text{Ca(OH)}_2) + (1 \text{ mol})(\Delta H_{\text{f}} \text{H}_2) - (1 \text{ mol})(\Delta H_{\text{f}} \text{Ca}) - (2 \text{ mol})(\Delta H_{\text{f}} \text{H}_2\text{O(l)})$$

$$\Delta H_{\text{rx}} = (1 \text{ mol})(-986.1 \text{ kJ/mol}) + (1 \text{ mol})(0 \text{ kJ/mol}) - (1 \text{ mol})(0 \text{ kJ/mol}) - (2 \text{ mol})(-285.8 \text{ kJ/mol})$$

$$\underline{\underline{\Delta H_{\text{rx}} = -414.5 \text{ kJ}}}$$

2) Calculate the amount of heat released when 15.0 g of methanol burns.



$$\Delta H_{\text{rxn}} = (1 \text{ mol})(\text{CO}_2\text{(g)}) + (2 \text{ mol})(2\text{H}_2\text{O(l)}) - (1 \text{ mol})(\text{CH}_3\text{OH(l)}) - (\frac{3}{2} \text{ mol})(\text{O}_2\text{(g)})$$

2

$$\Delta H_{\text{rxn}} = (1 \text{ mol})(-393.5 \text{ kJ/mol}) + (2 \text{ mol})(-285.8 \text{ kJ/mol}) - (1 \text{ mol})(-238.6 \text{ kJ/mol}) - (\frac{3}{2} \text{ mol})(0 \text{ kJ/mol})$$

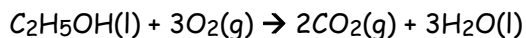
$$\Delta H_{\text{rxn}} = -726.5 \text{ kJ}$$

$$n = \frac{\text{mass}}{\text{molar mass}} = \frac{15.0 \text{ g}}{32.04216 \text{ g/mol}} = 0.468 \text{ mol methanol}$$

$$\frac{0.468 \text{ mol}}{1 \text{ mol}} = \frac{x}{-726.5 \text{ kJ}}$$

$$x = -340. \text{ kJ}$$

3) Calculate the mass of ethanol that must be burned in order to produce 20.0 kJ.



$$\Delta H_{\text{rxn}} = (2 \text{ mol})(\text{CO}_2\text{(g)}) + (3 \text{ mol})(\text{H}_2\text{O(l)}) - (1 \text{ mol})(\text{C}_2\text{H}_5\text{OH(l)}) - (3 \text{ mol})(\text{O}_2\text{(g)})$$

$$\Delta H_{\text{rxn}} = (2 \text{ mol})(-393.5 \text{ kJ/mol}) + (3 \text{ mol})(-285.8 \text{ kJ/mol}) - (1 \text{ mol})(-277.1 \text{ kJ/mol}) - (3 \text{ mol})(0 \text{ kJ/mol})$$

$$\Delta H_{\text{rxn}} = -1367.3 \text{ kJ}$$

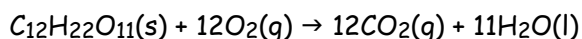
$$\frac{1 \text{ mol}}{x} = \frac{-1367.3 \text{ kJ}}{-20.0 \text{ kJ}}$$

$$x = 0.0146 \text{ mol}$$

$$\text{mass} = (\text{mol})(\text{mol mass}) = (0.0146 \text{ mol})(46.06904 \text{ g/mol}) = 0.674 \text{ g ethanol}$$

4) One teaspoon of sugar has a mass of 5.00g.

a) Calculate the energy change when a person consumes 2.00 teaspoons of sugar.



$$\Delta H_{rxn} = (12 \text{ mol})(CO_2(g)) + (11 \text{ mol})(H_2O(l)) - (1 \text{ mol})(C_{12}H_{22}O_{11}(s)) - (12 \text{ mol})(O_2(g))$$

$$\Delta H_{rxn} = (12 \text{ mol})(-393.5 \text{ kJ/mol}) + (11 \text{ mol})(-285.8 \text{ kJ/mol}) - (1 \text{ mol})(-2225.5 \text{ kJ/mol}) - (12 \text{ mol})(0 \text{ kJ/mol})$$

$$\Delta H_{rxn} = -5640.3 \text{ kJ}$$

$$n = \frac{\text{mass}}{\text{mol mass}} = \frac{10.0 \text{ g}}{342.30008 \text{ g/mol}} = 0.0292 \text{ mol}$$

$$\frac{1 \text{ mol}}{0.0292 \text{ mol}} = \frac{-5640.3 \text{ kJ}}{x}$$

$$x = -164.8 \text{ kJ}$$

b) Energy obtained from food is often reported in food calories, or Calories  
1 Calorie (1 Cal) = 1 kilocalorie (1 kcal) = 1000 calories (1000 cal) = 4.18 kJ  
Calculate the number of food calories contained in 2.00 teaspoons of sugar.

$$\begin{array}{r} = 4.18 \text{ kJ} \\ \times \quad 165 \text{ kJ} \\ \hline x = 39.4 \text{ Cal} \end{array}$$

c) A 150.0 lb person walking at a moderate speed will burn 40.0 Cal while walking for 15.0 minutes. How long would this person have to walk to burn the Calories in 2.00 teaspoons of sugar?  $\frac{1 \text{ Cal}}{15.0 \text{ min}}$

$$\begin{array}{r} = 40.0 \text{ Cal} \\ \times \quad 39.4 \text{ Cal} \\ \hline x = 14.8 \text{ minutes} \end{array}$$

5) Which sugar compound is more stable, glucose or sucrose? Explain.

$$\text{Glucose } \Delta H_f = -1273.1 \text{ kJ/mol}$$

$$\text{Sucrose } \Delta H_f = -2225.5 \text{ kJ/mol}$$

Both glucose and sucrose are very stable as evidenced by their large negative enthalpies of formation.

Sucrose is more stable than glucose based on the large negative enthalpy of formation.  $\frac{15.0 \text{ min}}{1 \text{ Cal}}$

- 6) What mass of hydrogen gas should be burned (under standard conditions) to heat 1.00 L of water from 10.0°C to 80.0°C?

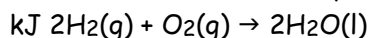
$$q = mc\Delta t$$

$$q = (1.00 \text{ kg})(4.18$$

$$\text{J/g}^\circ\text{C})(70.0^\circ\text{C}) \quad q = 292.6 \text{ kJ}$$

∴ the water requires 292.6 kJ

∴ the combustion of hydrogen must supply 292.6



$$\Delta H_{\text{rxn}} = (2 \text{ mol})(\text{H}_2\text{O}(\text{l})) - (2 \text{ mol})(2\text{H}_2(\text{g})) - (1 \text{ mol})(\text{O}_2(\text{g}))$$

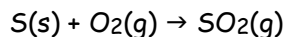
$$\Delta H_{\text{rxn}} = (2 \text{ mol})(-285.8 \text{ kJ/mol}) - (1 \text{ mol})(0 \text{ kJ/mol}) - (1 \text{ mol})(0 \text{ kJ/mol})$$

$$\Delta H_{\text{rxn}} = -571.8 \text{ kJ}$$

$$\begin{array}{l} = -571.8 \text{ kJ} \\ \times \quad -292.6 \text{ kJ} \end{array}$$

$$\text{mass} = (\text{mol})(\text{mol mass}) = (1.02 \text{ mol})(20.1588 \text{ g/mol}) = 2.06 \text{ g hydrogen}$$

- 7) What would be the final temperature of 885 g water, initially at 25.0°C, if it were heated by the burning of 20.0 g of sulfur in excess oxygen to produce sulfur dioxide?



$$\Delta H_{\text{rxn}} = (1 \text{ mol})(-296.8 \text{ kJ/mol}) - (1 \text{ mol})(0 \text{ kJ/mol}) - (1 \text{ mol})(0$$

$$\text{kJ/mol}) \quad \Delta H_{\text{rxn}} = -296.8 \text{ kJ}$$

$$n = \frac{\text{mass}}{\text{molmass}} = \frac{20.0\text{g}}{32.066\text{g/mol}} = 0.624 \text{ mol Sulfur}$$

$$\frac{1 \text{ mol}}{0.624 \text{ mol}} = \frac{-296.8 \text{ kJ}}{x}$$

$$x = -185 \text{ kJ}$$

∴ when 20.0 g of sulfur burns, 185 kJ is released

$$q = mc\Delta t$$

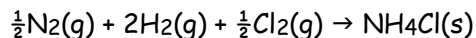
$$185 \text{ kJ} = (0.885 \text{ kg})(4.18 \text{ J/g}^\circ\text{C})(T - 25.0^\circ\text{C})$$

$$\left( \frac{185}{0.885 \text{ kg} \cdot 4.18 \text{ J/g}^\circ\text{C}} \right) = T - 25.0^\circ\text{C}$$

$$50.0^\circ\text{C} = T - 25.0^\circ\text{C}$$

$$T = 75.0^\circ\text{C}$$

- 8) When 4.00 g of ammonium chloride is formed from its elements, 23.5 kJ of energy is released.  
 a) Write the thermochemical equation showing the formation of ammonium chloride.



- b) Calculate the standard heat of formation for ammonium chloride.

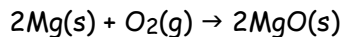
$$n = \frac{\text{mass}}{\text{mol mass}} = \frac{4.00 \text{ g}}{53.49146 \text{ g/mol}} = 0.0748 \text{ mol}$$

$$\frac{0.0748 \text{ mol}}{1 \text{ mol}} = \frac{-23.5 \text{ kJ}}{x}$$

$$x = -314 \text{ kJ}$$

$$\therefore \Delta H_f \text{ NH}_4\text{Cl}(\text{s}) = -314 \text{ kJ/mol}$$

- 9) What volume of water would be needed to absorb the energy released in burning 1.00 g of magnesium and show a temperature increase (of the water) of 10.0°C?



$$\Delta H_{\text{rxn}} = (2 \text{ mol})(601.7 \text{ kJ/mol}) - (2 \text{ mol})(0 \text{ kJ/mol}) - (1 \text{ mol})(0 \text{ kJ/mol})$$

$$\Delta H_{\text{rxn}} = -1203.4 \text{ kJ}$$

$$n = \frac{\text{mass}}{\text{mol mass}} = \frac{1.00 \text{ g}}{24.305 \text{ g/mol}} = 0.0411 \text{ mol}$$

$$\frac{2 \text{ mol}}{0.0411 \text{ mol}} = \frac{-1203.4 \text{ kJ}}{x}$$

$$x = -24.8 \text{ kJ}$$

∴ 24.8 kJ is released when 1.00 g of magnesium burns

∴ 24.8 kJ is absorbed by the water

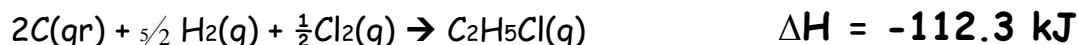
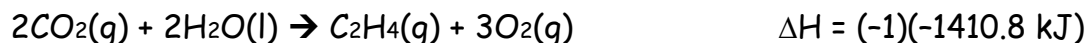
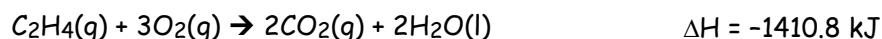
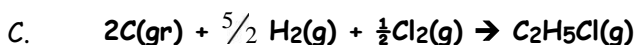
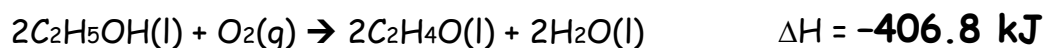
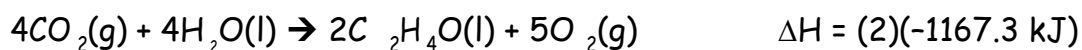
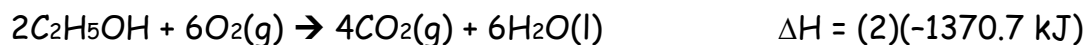
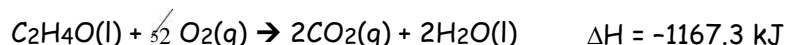
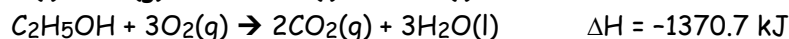
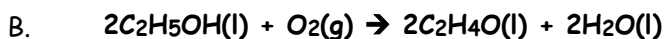
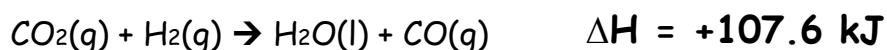
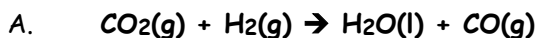
$$q = mc\Delta t$$

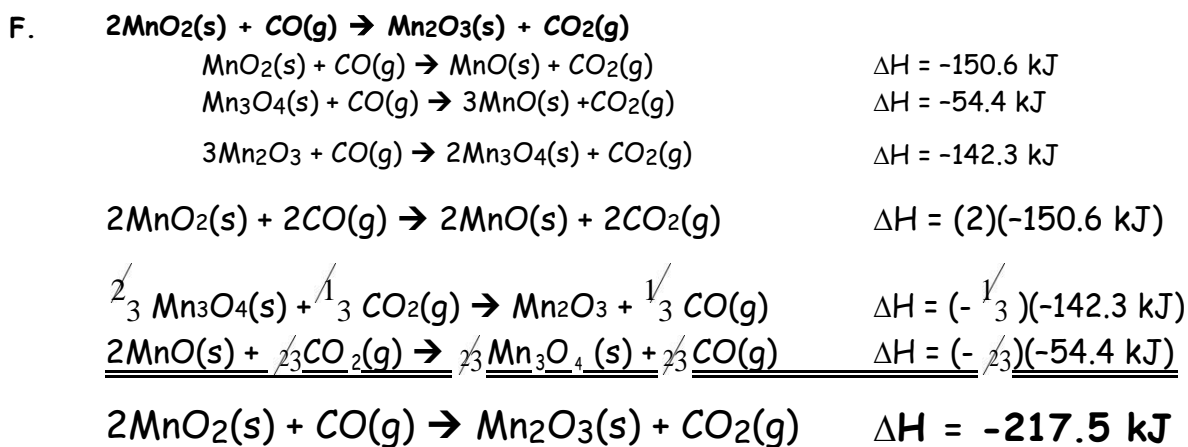
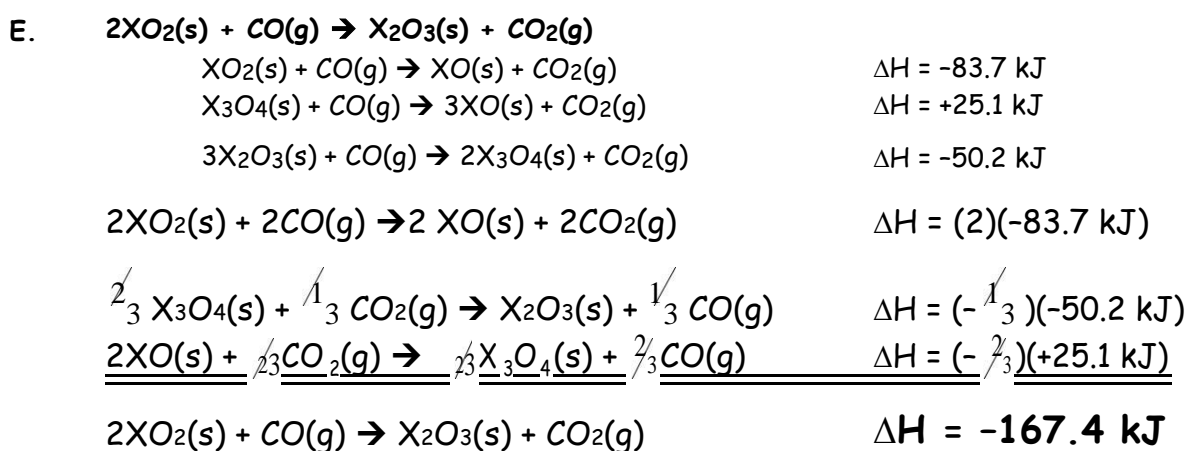
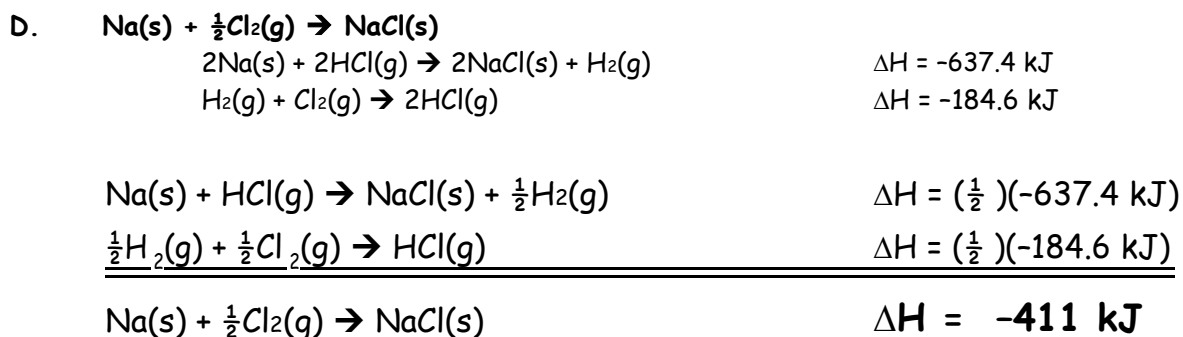
$$24.8 \text{ kJ} = m(4.18 \text{ kJ/kg}^\circ\text{C})(10.0^\circ\text{C})$$

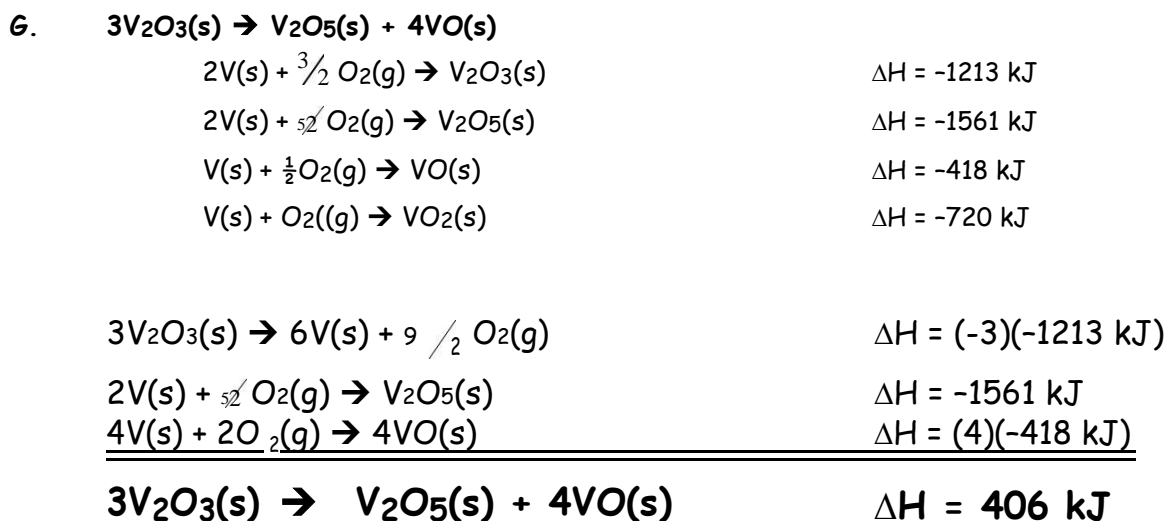
$$m = \left( \frac{24.8 \text{ kJ}}{4.18 \text{ kJ/kg}^\circ\text{C} \cdot 10.0^\circ\text{C}} \right) = 0.592 \text{ kg}$$

∴ 0.592 L or 592 mL of water is needed

# Hess' Law







Data as in G.

