1. Case A: 1 mol of helium is placed a rigid 5 L container with pressure 2 atm. The container is initially at is initially at 200 °C and is heated to 250 °C.

   Case B: 1 mol of helium is placed a flexible 5 L container with pressure 2 atm. The container is initially at is initially at 200 °C and is heated to 250 °C.

   (a) Which case has the higher q?

   A: Constant volume, $q = nC_V \Delta T$
   B: Constant pressure, $q = nC_P \Delta T$
   $C_P > C_V \implies q_B > q_A$

   (b) Which case has the higher ΔU?

   State function, same initial and final state, so same ΔU

2. 0.6192 g of sulfur is combusted in a calorimeter with a heat capacity of 5.270 kJ/°C. Both $SO_2(g)$ and $SO_3(g)$ are formed, as shown in the equations below. The temperature of the calorimeter increases by 1.140 °C. Determine the molar ratio of $SO_3(g)/SO_2(g)$ produced.

$$\mathrm{S\,(s)} + \mathrm{O_2(g)} \rightarrow \mathrm{SO_2(g)} \qquad \Delta H^\circ_{\mathrm{rxn}} = -296.83\,\mathrm{kJ}$$

$$\mathrm{S\,(s)} + \frac{3}{2}\mathrm{O_2(g)} \rightarrow \mathrm{SO_3(g)} \qquad \Delta H^\circ_{\mathrm{rxn}} = -395.72\,\mathrm{kJ}$$

Calculate initial mols of sulfur:

$$0.6192 \text{ g S} \left(\frac{1 \text{ mol}}{32.06 \text{ g}}\right) = 0.01931\,\text{mol S}$$

Of the 0.01931 mol of sulfur, the products can be either SO_2 or SO_3. If x mol of sulfur forms SO_2, then $n_{SO_2} = x$ and $n_{SO_3} = 0.01931 - x$

We write $q_{in} = -q_{out}$, where the enthalpies are multiplied by the mols and summed to get the heat produced, and the heat goes to increasing the temperature of the calorimeter.

$$\begin{aligned}
-\frac{296.83 \text{ kJ}}{\text{mol}} x - \frac{395.72 \text{ kJ}}{\text{mol}} (0.01931 - x) &= -C\Delta T \\
-296.83x + 395.72x - 7.6414 &= -\left(5.270 \text{ kJ/°C}\right)\left(1.140 \text{ °C}\right) \\
98.89x - 7.6414 &= -6.0068 \\
x &= 0.01653\,\text{mol}
\end{aligned}$$

$$\frac{n_{SO_3}}{n_{SO_2}} = \frac{0.01931 - x}{x} = \boxed{0.168}$$

3. Show that for a reversible adiabatic expansion of an ideal gas:

$$PV^\gamma = constant$$

Start from this result in lecture:

$$\frac{T_2}{T_1} = \left(\frac{V_1}{V_2}\right)^{\gamma-1}$$

Substitute the ideal gas law $T = \dfrac{PV}{nR}$:

$$\frac{P_2V_2}{P_1V_1} = \left(\frac{V_1}{V_2}\right)^{\gamma-1}$$

Cross multiply:

$$P_1V_1^\gamma = P_2V_2^\gamma$$

$$\boxed{PV^\gamma = \text{constant}}$$

4. The heat from the reaction

$$2\,\text{Al (s)} + 6\,\text{HBr (aq)} \longrightarrow 2\,\text{AlBr}_3\,\text{(aq)} + 3\,\text{H}_2\,\text{(g)} \qquad \Delta H^\circ_{\text{rxn}} = -1061\text{ kJ}$$

is completely transferred to water that is initially at 25 °C. How many grams of water can be converted to steam at 110 °C when 191 g of aluminum is added to excess HBr (aq)?

Quantity	Value
$\Delta H_{\text{vap, H}_2\text{O}}$	$2.261\,\frac{\text{kJ}}{\text{g}}$
$C_{s,\ \text{H}_2\text{O (l)}}$	$4.184\,\frac{\text{J}}{\text{g}\cdot{}^\circ\text{C}}$
$C_{s,\ \text{H}_2\text{O (g)}}$	$2.09\,\frac{\text{J}}{\text{g}\cdot{}^\circ\text{C}}$

Calculate heat released by reaction:

$$191\text{ g Al}\left(\frac{1\text{ mol}}{26.98\text{ g}}\right)\left(\frac{-1061\text{ kJ}}{2\text{ mol Al}}\right) = -3.756\times10^6\text{ J}$$

Water heating has 3 steps:

$$\begin{aligned} q_{\text{H}_2\text{O}} &= (q_{25\to100}) + (q_{\text{vap}}) + (q_{100\to110}) \\ &= m\left(4.184\,\tfrac{\text{J}}{\text{g}\cdot{}^\circ\text{C}}\right)(100-25)\,{}^\circ\text{C} + m\left(2.261\,\tfrac{\text{kJ}}{\text{g}}\right)\left(1{,}000\,\tfrac{\text{J}}{\text{kJ}}\right) + m\left(2.09\,\tfrac{\text{J}}{\text{g}\cdot{}^\circ\text{C}}\right)(110-100)\,{}^\circ\text{C}\end{aligned}$$

Setup conservation of energy:

$$q_{\text{in}} = -q_{\text{out}}$$

$$-3.756\times10^6\text{ J} = -(313.8m + 2261m + 20.9m)$$

$$\boxed{m = 1{,}447\,\text{g}}$$

Homework Problem 14

1. Show that for a reversible adiabatic expansion of an ideal gas:

$$\frac{T_2}{T_1} = \left(\frac{P_2}{P_1}\right)^{\frac{\gamma-1}{\gamma}}$$

Start from the adiabatic relation $P_1 V_1^\gamma = P_2 V_2^\gamma$