

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  $\Delta U$ ?

State function, same initial and final state, so same  $\Delta U$

2. 0.6192 g of sulfur is combusted in a calorimeter with a heat capacity of 5.270 kJ/°C. Both SO<sub>2</sub>(g) and SO<sub>3</sub>(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<sub>3</sub>(g)/SO<sub>2</sub>(g) produced.



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<sub>2</sub> or SO<sub>3</sub>. If  $x$  mol of sulfur forms SO<sub>2</sub>, then  $n_{\text{SO}_2} = x$  and  $n_{\text{SO}_3} = 0.01931 - x$

We write  $q_{\text{in}} = -q_{\text{out}}$ , 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.

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

$$\frac{n_{\text{SO}_3}}{n_{\text{SO}_2}} = \frac{0.01931 - x}{x} = \boxed{0.168}$$

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

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

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  $T = \frac{PV}{nR}$ :

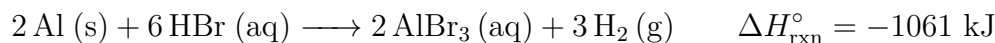
$$\frac{P_2V_2}{P_1V_1} = \left(\frac{V_1}{V_2}\right)^{\gamma-1}$$

Cross multiply:

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

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

4. The heat from the reaction



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:

$$191 \text{ g Al} \left(\frac{1 \text{ mol}}{26.98 \text{ g}}\right) \left(\frac{-1061 \text{ kJ}}{2 \text{ mol Al}}\right) = -3.756 \times 10^6 \text{ J}$$

Water heating has 3 steps:

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

Setup conservation of energy:

$$\begin{aligned} q_{\text{in}} &= -q_{\text{out}} \\ -3.756 \times 10^6 \text{ J} &= -(313.8m + 2261m + 20.9m) \end{aligned}$$

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

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