INTERNAL ENERGY PROBLEMS

PROBLEMS: (1 L-atm = 101.3 J)

1. In an exothermic process, the volume of a system expanded from 186 mL to 1997 mL against a constant pressure of 745 torr. During the process, 18.6 calories of heat energy were given off. What was the internal energy change for the system in joules?

$$q = -18.6 \text{ cal} \times \frac{4.184 \text{ J}}{1 \text{ cal}} = -77.8 \text{ J}$$

$$w = -P\Delta V = -745 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} \times (1997 \text{ mL} - 186 \text{ mL}) \times \frac{10^{-3} \text{ L}}{1 \text{ mL}} \times \frac{101.3 \text{ J}}{1 \text{ L atm}} = -17\underline{9}.8 \text{ J}$$

$$\Delta U = q + w = (-77.8 \text{ J}) + (-17\underline{9}.8 \text{ J}) = -258 \text{ J}$$

2. Calculate the change in internal energy for the thermal decomposition of 1.000 g of Potassium Chlorate at a constant external pressure of 943.2 mmHg. The decomposition reaction is

$$2\ KClO_3\ (s) \xrightarrow[MnO_2]{\Delta} 2\ KCl\ (s) + 3\ O_2\ (g)$$

Potassium Chlorate's heat of formation is -391.20 kJ/mol, Potassium Chloride's heat of formation is -435.87 kJ/mol, and Oxygen has a density of 1.308 g/L at the reaction temperature.

$$\Delta H = \left[3(0.00 \text{ kJ}) + 2(-435.87 \text{ kJ})\right] - \left[2(-391.20 \text{ kJ})\right] = -89.34 \text{ kJ / mol rxn}$$

$$1.000 \text{ g KClO}_3 \times \frac{1 \text{ mol KClO}_3}{122.5 \text{ g KClO}_3} \times \frac{1 \text{ mol rxn}}{2 \text{ mol KClO}_3} \times \frac{-89.34 \text{ kJ}}{1 \text{ mol rxn}} \times \frac{10^3 \text{ J}}{1 \text{ kJ}} = -364.5 \text{ J}$$

$$\Delta V = 1.000 \text{ g KClO}_3 \times \frac{1 \text{ mol KClO}_3}{122.5 \text{ g KClO}_3} \times \frac{3 \text{ mol O}_2}{2 \text{ mol KClO}_3} \times \frac{31.9988 \text{ g O}_2}{1 \text{ mol O}_2} \times \frac{1 \text{ L}}{1.308 \text{ g O}_2} = 0.2994 \text{ L}$$

$$\Delta U = \Delta H - P\Delta V = -364.5 \text{ J} - (943.2 \text{ mmHg})\left(\frac{1 \text{ atm}}{760 \text{ mmHg}}\right)(0.2994 \text{ L})\left(\frac{101.3 \text{ J}}{1 \text{ L atm}}\right) = -401.2 \text{ J}$$

3. (a) A gas at expands from 2.0 L to 6.0 L at a constant pressure of 912 mmHg. If q_p was zero in the process, what would be the change in internal energy?

$$\Delta U = q + w = 0.00\text{ J} - 921\text{ mmHg} \times \frac{1\text{ atm}}{760\text{ mmg}} \times (6.0\text{ L} - 2.0\text{ L}) \times \frac{101.3\text{ J}}{1\text{ L atm}} = -4.9 \times 10^{2}\text{ J}$$

(b) What would be the change in internal energy for the process described in 3 (a) if the expansion occurred when the external pressure was zero?

If the pressure is zero then the work is also zero. Therefore, the change in internal energy is zero.

4. The oxidation of Nitric Oxide

$$2\text{ NO }(g) + O_2\ (g) \rightarrow 2\text{ NO}_2\ (g) \qquad \Delta H^{\circ}_{rxn} = -113.1\text{ kJ}$$

is a key step in the production of photochemical smog. Calculate the change in internal energy (in kJ) that occurs when 15.4 g of NO reacts with excess Oxygen at 35.0°C.

$$\Delta U = \Delta H - \Delta nRT = -113.1\text{ kJ/mol rxn} - \left(\frac{-1\text{ mol gas}}{1\text{ mol rxn}}\right)\frac{8.314 \times 10^{-3}\text{ kJ}}{\text{mol gas K}}(308.2\text{ K})$$

$$= -110.5\text{ kJ/mol } O_2$$

$$q = 15.4\text{ g NO} \times \frac{1\text{ mol NO}}{30.0061\text{ g NO}} \times \frac{1\text{ mol rxn}}{2\text{ mol NO}} \times \frac{-110.5\text{ kJ}}{1\text{ mol rxn}} = -28.4\text{ kJ}$$

5. A gaseous mixture was enclosed in a piston and cylinder system having a volume of 1545 mL. When a chemical reaction occurred, the volume decreased to 375 mL and 321 calories of heat was absorbed. The external pressure was 753 mmHg. Calculate the change in enthalpy, work, and the change in internal energy for this system in kJ.

$$q_p = 321\text{ cal} \times \frac{4.184\text{ J}}{1\text{ cal}} \times \frac{1\text{ kJ}}{10^3\text{ J}} = 1.34\text{ kJ} = \Delta H$$

$$w = -P\Delta V = -753\text{ mmHg} \times \frac{1\text{ atm}}{760\text{ mmHg}} \times (375\text{ mL} - 1545\text{ mL}) \times \frac{10^{-3}\text{ L}}{1\text{ mL}} \times \frac{101.3\text{ J}}{1\text{ L atm}} \times \frac{1\text{ kJ}}{10^3\text{ J}} = 0.1174\text{ kJ}$$

$$\Delta U = q + w = 1.34\text{ kJ} + 0.117\text{ kJ} = 1.46\text{ kJ}$$

6. When 3.55 g of Methane (CH_4) gas, was burned in excess Oxygen at 45°C, the internal energy change was -196.3 kJ. Calculate the enthalpy change for the combustion of <u>one</u> mole of methane at 45°C.

$$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(l)$$

$$q_v = -196.3\text{ kJ}$$

$$\Delta U = \frac{q_v}{n} = \frac{-196.3\text{ kJ}}{3.55\text{ g }CH_4 \times \frac{1\text{ mol }CH_4}{16.04\text{ g }CH_4} \times \frac{1\text{ mol rxn}}{1\text{mol }CH_4}} = -88\underline{6}.9\text{ kJ/mol rxn}$$

$$\Delta H = \Delta U + \Delta nRT = -88\underline{6}.9\text{ kJ/mol rxn} + \left(\frac{-2\text{ mol gas}}{\text{mol rxn}}\right)\left(\frac{8.314\times10^{-3}\text{ kJ}}{\text{mol gas K}}\right)(318\text{ K}) = -892\text{ kJ/mol rxn}$$