

**Supplemental Questions  
for  
Thermodynamics 1**

1. How much energy is required to take 10.0 g of liquid mercury from 25.0 °C to its boiling point (356.6 °C) and then convert it completely to vapor at its boiling point? You are given that  $\Delta H_{vap}$  for mercury is +59.3 kJ/mole and that the heat capacity of liquid mercury is 27.983 J·mole<sup>-1</sup>·K<sup>-1</sup>. In your answer be sure to indicate the direction of heat flow in words (signs are not sufficient!).

*Question only asks about how much heat must be put into the system (the mercury). So, we can assume that all of the heat came from the surroundings and focus only on the system.*

*Mercury is warmed and then vaporized. Use the heat capacity for the heating portion (watch the units!) and  $\Delta H_{vap}$  for the change of state. Total energy required must be the sum of the energy for the two separate processes.*

$$q = nC\Delta T + n\Delta H_{vap}$$

$$q = 10.0g \left( \frac{1mole}{200.59g} \right) \left( 0.027983 \frac{kJ}{mole \cdot K} \right) ((356.6 - 25.0)K) + 10.0g \left( \frac{1mole}{200.59g} \right) \left( +59.3 \frac{kJ}{mole} \right)$$

$$q = (0.0498_{53} mole) \left( 0.027983 \frac{kJ}{mole \cdot K} \right) (+331.6K) + 2.95_{627} kJ$$

$$q = +0.462_{593} kJ + 2.95_{627} kJ = +3.42kJ$$

**To convert 10.0 g from a liquid at 25.0 °C to a vapor at its boiling point requires that 3.42 kJ of energy as heat be added to the liquid.**

2. When one mole of CO<sub>2</sub> (s) is converted to CO<sub>2</sub> (g) at atmospheric pressure and -78 °C, the heat absorbed by the system exceeds the increase in the internal energy of the CO<sub>2</sub>. Explain what happened to the “missing” energy.

**Since  $q > \Delta U$ ,  $\Delta U - q < 0$ , which means that  $w$  is also less than zero (from the First Law of Thermodynamics). Because  $w$  is negative, we know that the system did work. So, the “missing” energy was work that the system did (FYI the work was done when the CO<sub>2</sub> (g) pushed the atmosphere out of the way).**

3a. Calculate  $\Delta H^0$  for the reaction  $\text{N}_2\text{H}_4(l) + \text{O}_2(g) \rightarrow \text{N}_2(g) + 2 \text{H}_2\text{O}(l)$  from the reactions given below.

	$\Delta H^0$ (kJ)
$2 \text{NH}_3(g) + 3 \text{N}_2\text{O}(g) \rightarrow 4 \text{N}_2(g) + 3 \text{H}_2\text{O}(l)$	-1011.4
$\text{N}_2\text{O}(g) + 3 \text{H}_2(g) \rightarrow \text{N}_2\text{H}_4(l) + \text{H}_2\text{O}(l)$	-317.25
$2 \text{NH}_3(g) + 1/2 \text{O}_2(g) \rightarrow \text{N}_2\text{H}_4(l) + \text{H}_2\text{O}(l)$	-142.98
$\text{H}_2(g) + 1/2 \text{O}_2(g) \rightarrow \text{H}_2\text{O}(l)$	-285.83

*We want the reaction  $\text{N}_2\text{H}_4(l) + \text{O}_2(g) \rightarrow \text{N}_2(g) + 2 \text{H}_2\text{O}(l)$ , but it turns out that it is easier to get to the  $\Delta H^0$  reaction  $4 \text{N}_2\text{H}_4(l) + 4 \text{O}_2(g) \rightarrow 4 \text{N}_2(g) + 8 \text{H}_2\text{O}(l)$ .*

	$\Delta H^0$ (kJ)
$2 \text{NH}_3(g) + 3 \text{N}_2\text{O}(g) \rightarrow 4 \text{N}_2(g) + 3 \text{H}_2\text{O}(l)$	-1011.4
$3 \text{N}_2\text{H}_4(l) + 3 \text{H}_2\text{O}(l) \rightarrow 3 \text{N}_2\text{O}(g) + 9 \text{H}_2(g)$	$3(+317.25) = +951.75$
$\text{N}_2\text{H}_4(l) + \text{H}_2\text{O}(l) \rightarrow 2 \text{NH}_3(g) + 1/2 \text{O}_2(g)$	+142.98
$9 \text{H}_2(g) + 9/2 \text{O}_2(g) \rightarrow 9 \text{H}_2\text{O}(l)$	$9(-285.83) = -2572.47$
$4 \text{N}_2\text{H}_4(l) + 4 \text{O}_2(g) \rightarrow 4 \text{N}_2(g) + 8 \text{H}_2\text{O}(l)$	-2489.1

Since this reaction is four times the one we want, divide everything by 4.

$\Delta H^0$  for this reaction is -622.29 kJ

b. If 10.0 g of hydrazine,  $\text{N}_2\text{H}_4$ , is allowed to react with an excess amount of  $\text{O}_2$ , how much heat would be exchanged with the surroundings? In your answer clearly state the direction of the heat flow. The molar mass of  $\text{N}_2\text{H}_4$  is 32.05 g/mole.

*There is a chemical change involved. So, start with the equation for heat of a chemical change.*

$$q_{\text{rxn}} = n \Delta H$$

$$q = 10.0 \text{ g } \text{N}_2\text{H}_4 \left( \frac{1 \text{ mole } \text{N}_2\text{H}_4}{32.05 \text{ g } \text{N}_2\text{H}_4} \right) \left( \frac{-622.29 \text{ kJ}}{1 \text{ mole } \text{N}_2\text{H}_4} \right) = -194.16 \text{ kJ}$$

In this reaction 194 kJ of heat are released to the surroundings.

c. If all of the heat from part b is transferred to a 500.0 g sample of solid water at 0.00 °C, how what will be the final temperature of the water? The specific heat capacity of liquid water is 4.184 J·g<sup>-1</sup>·K<sup>-1</sup>, the specific heat capacity of solid water is 2.06 J·g<sup>-1</sup>·K<sup>-1</sup>,  $\Delta H_{fus}$  for water is +6.008 kJ/mole, and the molar mass of water is 18.02 g/mole.

*We expect that the solid water will begin to melt, so we will use  $\Delta H_{fus}$ . But it is also possible that we might completely melt ice, and then begin to warm up the liquid water. So, we'll need to include both in our expression and be careful as we solve the problem.*

$$q_{rxn} = -q_{H_2O} = -(n_{H_2O}\Delta H_{fus} + C_{H_2O}m_{H_2O}\Delta T_{H_2O}) = -194.16 \text{ kJ}$$

$$-\left(500.0 \text{ g} \left(\frac{1 \text{ mole}}{18.02 \text{ g}}\right) (+6.008 \text{ kJ} \cdot \text{mole}^{-1}) + (4.184 \text{ J} \cdot \text{g}^{-1} \cdot \text{K}^{-1})(500.0 \text{ g})\Delta T_{H_2O}\right) = -194.16 \text{ kJ}$$

$$-((+166.7 \text{ kJ}) + (2.092 \text{ kJ} \cdot \text{K}^{-1})\Delta T_{H_2O}) = -194.16 \text{ kJ}$$

$$-166.7 \text{ kJ} + (-2.092 \text{ kJ} \cdot \text{K}^{-1})\Delta T_{H_2O} = -194.16 \text{ kJ}$$

$$(-2.092 \text{ kJ} \cdot \text{K}^{-1})\Delta T_{H_2O} = -27.46 \text{ kJ}$$

$$\Delta T_{H_2O} = +13. \text{ K}$$

**The final temperature of the water is 13. °C.**

*Note that if we didn't have enough heat to melt all of the ice, we would have gotten a nonsensical answer for  $\Delta T$ . We would have then known that we did not melt all of the ice and that the final temperature was 0.00 °C*

4a. The enthalpy change associated with KOH (s) dissolving in water is -57.61 kJ/mole. What will be the final temperature of a solution prepared by dissolving 100.0 g KOH in 1000.0 g of water that is initially at 20.0 °C? The formula weight of KOH is 56.11 g/mole. Assume that the specific heat capacity of the solution is 4.184 J·g<sup>-1</sup>·K<sup>-1</sup>.

*Assuming that all heat that the reaction, KOH (s) → KOH (aq), generates, or takes in, is transferred to, or from, the solution, we can write*

$$q_{rxn} = -q_{solution}$$

*The KOH undergoes a physical change (dissolving) and the solution undergoes a heating or cooling. The above equation can then be written as*

$$n \cdot \Delta H_{rxn} = -C_{solution} \cdot m_{solution} \cdot \Delta T_{solution}$$

*Substitute in known quantities (note that mass of solution includes the KOH) to give*

$$100.0 \text{ g KOH} \left( \frac{1 \text{ mole KOH}}{56.11 \text{ g KOH}} \right) \left( \frac{-57.61 \text{ kJ}}{1 \text{ mole}} \right) \left( \frac{1000 \text{ J}}{1 \text{ kJ}} \right) = - \left( 4.184 \frac{\text{J}}{\text{g} \cdot \text{K}} \right) (1100.0 \text{ g}) \Delta T_{solution}$$

$$-102.67 \times 10^3 \text{ J} = -4602.4 \frac{\text{J}}{\text{K}} \Delta T_{solution}$$

$$\Delta T_{solution} = +22.31 \text{ K} = +22.31 \text{ }^\circ\text{C}$$

$$\Delta T_{solution} = T_{final} - T_{initial} = T_{final} - 20.00 \text{ }^\circ\text{C} = +22.31 \text{ }^\circ\text{C}$$

$$T_{final} = 42.31 \text{ }^\circ\text{C}$$

**The final temperature of the solution is 42.3 °C.**

b. What is the molarity of this solution? Assume that the density of the solution is 1.00 g/mL.

$$[\text{KOH}] = \frac{1.782 \text{ mole KOH}}{1100.0 \text{ g solution}} \left( \frac{1.00 \text{ g solution}}{1 \text{ mL solution}} \right) \left( \frac{1000 \text{ mL}}{1 \text{ L}} \right) = 1.62 \text{ M}$$

**The concentration of KOH in this solution is 1.62 M.**

c. What practical implications does this have for the preparation of concentrated KOH solutions?

**When preparing concentrated KOH solutions it is advised that one take care because they are going to get hot.**

5a.  $\Delta H_{rxn}^0$  for the dissolution of  $\text{NH}_4\text{NO}_3$  in water is +25.69 kJ/mole. If 25.00 g  $\text{NH}_4\text{NO}_3$  is dissolved in 100.00 g of water (originally at 25.0 °C), what will be the temperature of the solution? Assume that the specific heat capacity of the solution is  $4.184 \text{ J} \cdot \text{g}^{-1} \cdot \text{K}^{-1}$ , and the density of the solution is  $1.000 \text{ g/cm}^3$ . The molar mass of  $\text{NH}_4\text{NO}_3$  is 80.04 g/mole.

*Again, assume that all heat from the system is transferred to, or from, the surroundings. And then realize that the  $\text{NH}_4\text{NO}_3$  undergoes a physical change, but the solution heats up/cool down.*

$$q_{rxn} = -q_{solution}$$

$$n\Delta H_{rxn} = -C_{solution}m_{solution}\Delta T_{solution}$$

$$25.00 \text{ g} \left( \frac{1 \text{ mole}}{80.04 \text{ g}} \right) \left( \frac{+25.69 \text{ kJ}}{1 \text{ mole}} \right) = -(4.184 \text{ J} \cdot \text{g}^{-1} \cdot \text{K}^{-1})(125.00 \text{ g})\Delta T_{solution}$$

$$+8.024_1 \text{ kJ} = -(0.5230 \text{ kJ} \cdot \text{K}^{-1})\Delta T_{solution}$$

$$\Delta T_{solution} = -15.3 \text{ K} = -15.3 \text{ }^\circ\text{C}$$

$$\Delta T_{solution} = T_{final} - T_{initial} = T_{final} - 25.0 \text{ }^\circ\text{C} = -15.3 \text{ }^\circ\text{C}$$

$$T_{final} = 9.7 \text{ }^\circ\text{C}$$

**The final temperature of the solution is 9.7 °C.**

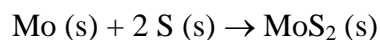
b. This “reaction” is used commercially in cold packs. In these cold packs the amount of  $\text{NH}_4\text{NO}_3$  is calculated to give a saturated solution. Explain why.

**If the solution is saturated, the maximum amount of  $\text{NH}_4\text{NO}_3$  will dissolve, and the maximum amount of heat will be absorbed.**

6. Molybdenum(IV) sulfide,  $\text{MoS}_2$ , is used in several important industrial processes. From the data given below, calculate  $\Delta H_f^\circ$  ( $\text{MoS}_2$ , s), the standard molar enthalpy of formation for  $\text{MoS}_2$ . Show all work.

	$\Delta H^\circ$ (kJ)
$\text{MoS}_2$ (s) + 2 $\text{O}_3$ (g) $\rightarrow$ $\text{MoO}_2$ (s) + 2 $\text{SO}_2$ (g)	-1231.4
$\text{Mo}$ (s) + $\text{O}_2$ (g) $\rightarrow$ $\text{MoO}_2$ (s)	-588.5
$\text{Mo}_2\text{S}_3$ (s) + $\text{S}$ (s) $\rightarrow$ 2 $\text{MoS}_2$ (s)	-83.2
$\text{Mo}_2\text{S}_3$ + 3 $\text{O}_2$ (g) $\rightarrow$ 2 $\text{Mo}$ (s) + 3 $\text{SO}_2$ (g)	-502.4
2 $\text{Mo}$ (s) + 3 $\text{S}$ (s) $\rightarrow$ $\text{Mo}_2\text{S}_3$ (s)	-386.2

**Reaction that we need is**



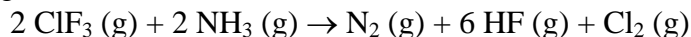
*Look for reactions that have  $\text{MoS}_2$  (s) as either a reactant or a product and arrange them so that they add up to the desired reaction.*

	$\Delta H^\circ$ (kJ)
$\text{Mo}_2\text{S}_3$ (s) + $\text{S}$ (s) $\rightarrow$ 2 $\text{MoS}_2$ (s)	-83.2
<u>2 <math>\text{Mo}</math> (s) + 3 <math>\text{S}</math> (s) <math>\rightarrow</math> <math>\text{Mo}_2\text{S}_3</math> (s)</u>	<u>-386.2</u>
2 $\text{Mo}$ (s) + 4 $\text{S}$ (s) $\rightarrow$ 2 $\text{MoS}_2$ (s)	-469.4

**This is twice the desired reaction, so divide the reaction and  $\Delta H^\circ$  by 2 (moles).**

**$\Delta H_f^\circ$  ( $\text{MoS}_2$ , s) is -234.7 kJ/mole.**

7a. The following reaction has a  $\Delta H^0$  of -1196. kJ. Use this information and a table of  $\Delta H_f^0$  to calculate  $\Delta H_f^0$  ( $\text{ClF}_3$ , g).



*Enthalpy of reaction will be the enthalpies of the products minus the enthalpies of the reactants. So,*

$$\Delta H^0 = \Delta H_f^0 (\text{N}_2, \text{g}) + 6 \Delta H_f^0 (\text{HF}, \text{g}) + \Delta H_f^0 (\text{Cl}_2, \text{g}) - 2 \Delta H_f^0 (\text{ClF}_3, \text{g}) - 2 \Delta H_f^0 (\text{NH}_3, \text{g})$$

$$-1196. \text{kJ} = 0 + (6 \text{ mole})(-271.1 \text{ kJ/mole}) + 0 - 2 \Delta H_f^0 (\text{ClF}_3, \text{g}) - (2 \text{ mole})(-46.1 \text{ kJ/mole})$$

$$-1196. \text{kJ} = -1626.6 \text{ kJ} + (+92.2 \text{ kJ}) - (2 \text{ mole})\Delta H_f^0 (\text{ClF}_3, \text{g})$$

$$\Delta H_f^0 (\text{ClF}_3, \text{g}) = -169. \text{kJ/mole}$$

**$\Delta H_f^0$  for  $\text{ClF}_3$  gas is -169. kJ/mole.**

b. A 5.00-L vessel at 25.0 °C is filled with the reactants until the partial pressure of  $\text{ClF}_3$  is 325.0 mm Hg and the partial pressure of  $\text{NH}_3$  is 401.0 mm Hg. The temperature is raised to 277.3 °C and is held there as the reaction occurs. What is the total pressure in the vessel when the reaction is over?

*First find moles of each reactant, assuming that they behave ideally at the given temperature and pressure.*

$$n_{\text{ClF}_3} = \frac{PV}{RT} = \frac{\left(325.0 \text{ mm Hg} \left(\frac{1 \text{ atm}}{760 \text{ mm Hg}}\right)\right)(5.00 \text{ L})}{(0.082057 \text{ L} \cdot \text{atm} \cdot \text{K}^{-1} \cdot \text{mole}^{-1})(298.15 \text{ K})} = 0.0874_0 \text{ mole ClF}_3$$

$$n_{\text{NH}_3} = \frac{PV}{RT} = \frac{\left(401.0 \text{ mm Hg} \left(\frac{1 \text{ atm}}{760 \text{ mm Hg}}\right)\right)(5.00 \text{ L})}{(0.082057 \text{ L} \cdot \text{atm} \cdot \text{K}^{-1} \cdot \text{mole}^{-1})(298.15 \text{ K})} = 0.107_8 \text{ mole NH}_3$$

**Since 2 moles of  $\text{ClF}_3$  reacts with 2 moles of  $\text{NH}_3$ , we can see by inspection that  $\text{ClF}_3$  is the limiting reagent.** *We don't need to go all the way to moles of  $\text{ClF}_3$ , because we will need to know if there were any of the gaseous reactants will be left over.*

If 0.0874<sub>0</sub> mole ClF<sub>3</sub> reacts completely, then 0.0874<sub>0</sub> mole NH<sub>3</sub> will also react, and there will be 0.020<sub>4</sub> mole NH<sub>3</sub> remaining.

$$0.0874_0 \text{ mole ClF}_3 \left( \frac{1 \text{ mole N}_2}{2 \text{ mole ClF}_3} \right) = 0.0437_0 \text{ mole N}_2$$

$$0.0874_0 \text{ mole ClF}_3 \left( \frac{6 \text{ mole HF}}{2 \text{ mole ClF}_3} \right) = 0.262_2 \text{ mole HF}$$

$$0.0874_0 \text{ mole ClF}_3 \left( \frac{1 \text{ mole Cl}_2}{2 \text{ mole ClF}_3} \right) = 0.0437_0 \text{ mole Cl}_2$$

The total number of moles of gas remaining in the flask once the reaction is over is thus

$$0.0204 \text{ mole NH}_3 + 0.0437_0 \text{ mole N}_2 + 0.262_2 \text{ mole HF} + 0.0437_0 \text{ mole Cl}_2 = 0.370 \text{ mole}$$

The total pressure will be the sum of the partial pressures of each gas, which will be the same as using the total number of moles of gas in the Ideal Gas Law.

$$P = \frac{nRT}{V} = \frac{(0.370 \text{ mole})(0.082057 \text{ L} \cdot \text{atm} \cdot \text{K}^{-1} \cdot \text{mole}^{-1})(550.4_5 \text{ K})}{(5.00 \text{ L})} = 3.34 \text{ atm}$$

The final pressure in the flask is 3.34 atm.

c. You could have *qualitatively* predicted the change in pressure from only the balanced chemical reaction in part a. Explain.

**Four moles of gaseous reactants give eight moles of gaseous products, a net gain of four moles of gas. As the amount of gas increases at constant temperature and volume, the pressure must rise as dictated by the Ideal Gas Law.**

d. Why might this reaction be particularly hazardous to carry out?

**This is an exothermic reaction that generates gas and as such is an explosion hazard. If all of the heat given off by the reaction were used to heat the product gases at constant volume, it would cause the pressure inside the flask to rise. This increase in pressure coupled with the pressure increase described above could raise the pressure in the flask above its bursting point.**

8. In the production of cement calcium carbonate,  $\text{CaCO}_3$ , must be converted to calcium oxide,  $\text{CaO}$ , according to the following balanced chemical equation.



a. From a table of standard molar enthalpies of formation, calculate  $\Delta H_{\text{rxn}}^0$  for this reaction.

$$\Delta H_{\text{rxn}}^0 = (1 \text{ mole})(\Delta H_f^0 (\text{CaO}, \text{s})) + (1 \text{ mole})(\Delta H_f^0 (\text{CO}_2, \text{g})) - (1 \text{ mole})(\Delta H_f^0 (\text{CaCO}_3, \text{s}))$$

$$\Delta H_{\text{rxn}}^0 = (1 \text{ mole})(-635.1 \text{ kJ/mole}) + (1 \text{ mole})(-303.5 \text{ kJ/mole}) - (1 \text{ mole})(-1206.9 \text{ kJ/mole})$$

$$\Delta H_{\text{rxn}}^0 = +178.3 \text{ kJ}$$

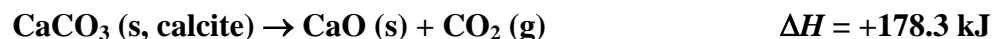
**$\Delta H_{\text{rxn}}^0$  for this reaction is +178.3 kJ**

b. Is this an endothermic or exothermic reaction?

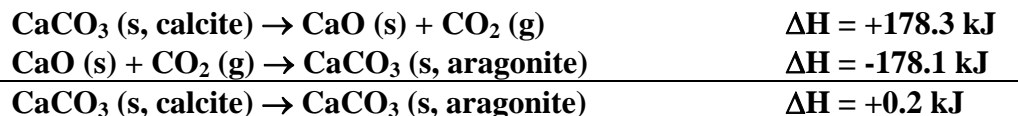
**This is an endothermic reaction since  $\Delta H$  is positive.**

d.  $\text{CaCO}_3$  actually exists in two forms. The calculation you did in part a was for the conversion of the form of  $\text{CaCO}_3$  known as calcite to  $\text{CaO}$ . If  $\Delta H_{\text{rxn}}$  for the conversion of aragonite (the other form of  $\text{CaCO}_3$ ) to  $\text{CaO}$  is +178.1 kJ/mole, what is  $\Delta H$  for the conversion of calcite to aragonite?

*The reactions that we have (written per 1 mole of the different  $\text{CaCO}_3$  forms) are*

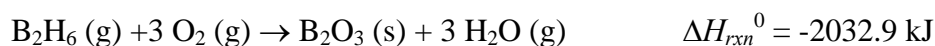


*We want  $\Delta H$  for  $\text{CaCO}_3 (\text{s, calcite}) \rightarrow \text{CaCO}_3 (\text{s, aragonite})$ , but reversing the second reaction and adding it to the first gives us the desired equation. Hess's Law says that if we then add up the  $\Delta H$  values we will have the desired  $\Delta H$ .*



**$\Delta H$  for the conversion of calcite to aragonite is +0.2 kJ.**

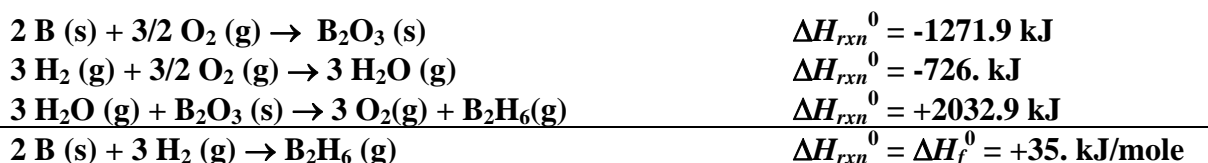
9. The standard molar enthalpy of formation,  $\Delta H_f^0$ , for diborane,  $B_2H_6(g)$ , cannot be found directly because the compound cannot be prepared by the reaction of boron and hydrogen. It can be calculated, however, using the following reactions:



a. Calculate  $\Delta H_f^0$  for  $B_2H_6(g)$ .

**Need the reaction which defined  $\Delta H_f^0(B_2H_6,g)$ :  $2 B(s) + 3 H_2(g) \rightarrow B_2H_6(g)$**

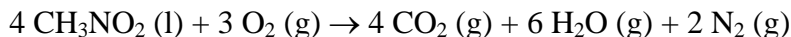
*Since we want 2 B on the reactant side, divide Eqn 1. by 2. If we reverse Eqn. 3, that will remove the  $B_2O_3(s)$  that we have as a product in Eqn. 1. Multiplying Eqn. 2 by 3/2 lets us remove the  $H_2(g)$  that we don't need. Summing up*



b. Would the calculated  $\Delta H_f^0$  for  $B_2H_6(g)$  change if  $H_2O(g)$  were replaced by  $H_2O(l)$  in the above equations? Why?

**No, it doesn't matter because enthalpy is a state function.  $\Delta H$  is independent of the path taken, so it doesn't matter if both  $H_2O$  are in the gaseous or liquid state. It also doesn't matter if we change one of the  $H_2O$ 's to the gaseous state as long as we take into account  $\Delta H_{vap}^0$  for  $H_2O$ .**

10. Nitromethane,  $\text{CH}_3\text{NO}_2$ , is used as a fuel in rockets and drag racers. It reacts with oxygen according to the following balanced chemical equation.



a. Use a table of  $\Delta H_f^\circ$  to calculate  $\Delta H^\circ$  for this reaction, given that  $\Delta H_f^\circ$  for liquid nitromethane is  $-113.1 \text{ kJ/mole}$ .

$$\Delta H^\circ = 4 \Delta H_f^\circ(\text{CO}_2, \text{g}) + 6 \Delta H_f^\circ(\text{H}_2\text{O}, \text{g}) - 4 \Delta H_f^\circ(\text{CH}_3\text{NO}_2, \text{l})$$

$$\Delta H^\circ = (4 \text{ mole})(-393.5 \text{ kJ/mole}) + (6 \text{ mole})(-241.8 \text{ kJ/mole}) - (4 \text{ mole})(-113.1 \text{ kJ/mole})$$

$$\Delta H^\circ = -1574. \text{ kJ} + -1450.8 \text{ kJ} + 452.4 \text{ kJ}$$

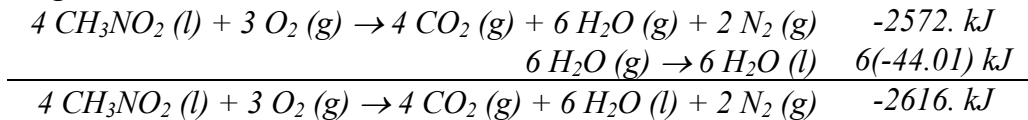
$$\Delta H^\circ = -2572. \text{ kJ}$$

**$\Delta H^\circ$  for this reaction is  $-2572. \text{ kJ}$**

b. Fuel cells were developed by NASA for use on spacecraft. In a fuel cell, a fuel reacts with oxygen without a flame and so the reaction takes place at a much lower temperature than the corresponding combustion reaction. If nitromethane is allowed to react with  $\text{O}_2$  in a fuel cell operating at  $25.0^\circ\text{C}$ , will  $\Delta H$  for the process be greater than, less than or the same as when nitromethane is burned? Explain.

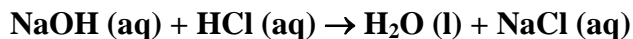
**The reaction in the fuel cell takes place at a much lower temperature which means that the  $\text{H}_2\text{O}$  that is the product of the reaction is not converted into a gas. The heat that went into the change of state from a liquid to a gas is not included in  $\Delta H$ , so the magnitude of  $\Delta H$  will be larger.**

*Using Hess's Law*

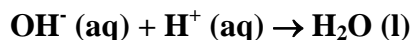


11. Aqueous sodium hydroxide will react with aqueous hydrochloric acid in an acid-base reaction. This reaction has a  $\Delta H^0$  of -55.84 kJ.

a. Write the balanced chemical equation for this reaction.



b. Write the net ionic equation for this reaction.



c. If 25.0 mL of a 2.15 M hydrochloric solution is mixed with 25.0 mL of a 2.00 M sodium hydroxide solution (both solutions at 23.4 °C initially), what will be the temperature of the solution when the reaction is over? Assume that the specific heat of all solutions is  $4.184 \text{ J}\cdot\text{g}^{-1}\cdot\text{K}^{-1}$ , and that the density of all solutions is  $1.00 \text{ g/cm}^3$ .

*Determine limiting reagent. We don't really need to calculate moles, or grams, of  $\text{H}_2\text{O}$ , because we will need to have moles of whichever reactant was the limiting reagent for the next calculation.*

$$25.0 \times 10^{-3} \text{ L} \left( \frac{2.15 \text{ mole HCl}}{1 \text{ L}} \right) = 0.0537_5 \text{ mole HCl}$$

$$25.0 \times 10^{-3} \text{ L} \left( \frac{2.00 \text{ mole NaOH}}{1 \text{ L}} \right) = 0.0500 \text{ mole NaOH}$$

**Since the stoichiometry is 1:1, it is clear by inspection that NaOH is the limiting reagent.**

**The densities of the solutions are  $1.00 \text{ g/cm}^3$ , so the mass of each is 25.0 g, and the total mass is 50.0 g.**

$$q_{\text{rxn}} = -q_{\text{solution}}$$

$$n\Delta H = -Cm\Delta T$$

$$(0.0500 \text{ mole}) \left( -55.84 \frac{\text{kJ}}{\text{mole}} \right) = - \left( 4.184 \frac{\text{J}}{\text{g}\cdot\text{K}} \right) (50.0 \text{ g})(\Delta T)$$

$$-2.79_2 \text{ kJ} = - \left( 0.209_2 \frac{\text{kJ}}{\text{K}} \right) (\Delta T)$$

$$\Delta T = \frac{-2.79_2 \text{ kJ}}{-0.209_2 \frac{\text{kJ}}{\text{K}}} = +13.3\text{K} = +13.3^\circ\text{C}$$

$$\Delta T = T_f - T_i = +13.3^\circ\text{C} = T_f - 23.1^\circ\text{C}$$

$$T_f = 36.7\text{ }^\circ\text{C}.$$

**The final temperature of the solution is 36.7 °C.**

d. When this experiment is actually done, the amount of heat measured will be less than what was calculated in part c. This is not a gross error! Where does the heat go? How do we correct for this?

**The real system is leaky. Some of the heat given off by the reaction leaves the solution, and is not accounted for in the measured  $\Delta T$ . We correct for this by running a reaction of known  $\Delta H$  in our reaction vessel (calorimeter) and calculating a specific heat capacity for the solution and the calorimeter (the calorimeter constant).**

e. The  $\Delta H^0$  for this reaction we only need to know  $\Delta H_f^0(\text{H}^+, \text{aq})$ ,  $\Delta H_f^0(\text{OH}^-, \text{aq})$  and  $\Delta H_f^0(\text{H}_2\text{O}, \text{l})$ . Explain why.

**In this reaction  $\text{Na}^+$  and  $\text{Cl}^-$  are spectator ions, so they do not change upon going from the reactant to the products (they are  $\text{Na}^+(\text{aq})$  and  $\text{Cl}^-(\text{aq})$  on both the reactant and product side of the chemical equation). Because they don't change during the course of the reaction, their  $\Delta H_f^0$  would appear on both sides of the expression for  $\Delta H_{rxn}$ , and would thus cancel out. The only things that change are  $\text{H}^+$ ,  $\text{OH}^-$  and  $\text{H}_2\text{O}$ , so only their  $\Delta H_f^0$  contribute to  $\Delta H_{rxn}$ .**

12. In tables you will find listed  $\Delta H_f^0(\text{H}_2\text{O}, \text{l})$  and  $\Delta H_f^0(\text{H}_2\text{O}, \text{g})$ , from which you can calculate  $\Delta H_{\text{vap}}^0$  for  $\text{H}_2\text{O}$  of +44.01 kJ/mole. In the same tables you will find  $\Delta H_{\text{vap}}$  of  $\text{H}_2\text{O}$  listed as +40.7 kJ/mole. Why are these numbers different? (Hint: it has to do with the definitions of the symbols).

**They are measured under different conditions.  $\Delta H_{\text{vap}}^0$  is measured under standard conditions (1 bar of pressure, 25.0 °C), while  $\Delta H_{\text{vap}}$  is measured at 100.0 °C (and 1 atm of pressure, although you don't know this, and weren't expect to specify pressure). Values of  $\Delta H$  for the same process measured under different conditions are not required to be the same, and often are not.**