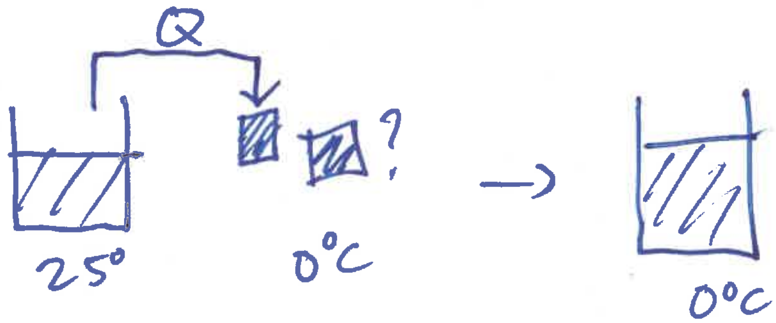
CHEM 101A – Chapter 6 Enthalpy and Stoichiometry

1. What mass of ice would be required to cool 355 grams of water from 25°C to 0°C? (specific heat capacity water = 4.184 J/g·°C and ΔH_{fusion} for ice = 6.020 kJ/mol.)

Assume ice = 0°C

$$H_2O(s) \rightarrow H_2O(l) \quad \Delta H = 6.020\ kJ/mol$$



$$Q_{H_2O(cool)} = -Q_{ice\text{-}melting}$$

$$Q = mC\Delta T$$

$$= (355\cancel{g})(4.184\ \frac{J}{\cancel{g}\cdot\cancel{°C}})(-25\cancel{°C})$$

$$= -37{,}133\ J \qquad (-37\ kJ)$$

$$-37\ kJ = -Q_{ice\text{-}melting}$$

$$Q_{ice\text{-}melting} = 37\ \cancel{kJ} \times \frac{1\ mol}{6.020\ \cancel{kJ}} \times \frac{18.02\ g}{1\ mol}$$

$$= \boxed{110\ g\ ice}$$

2. Consider the acid-base reaction below:

$$2\,HCl(aq) + Ba(OH)_2(aq) \rightarrow BaCl_2(aq) + 2H_2O(l) \quad \Delta H = -118\text{ kJ}$$

a. What quantity of heat is released when 100.0 mL 0.500 *M* HCl are mixed with 300.0 mL 0.100 *M* $Ba(OH)_2(aq)$.

$\frac{0.500\text{ mol HCl}}{L} \times 0.1000L = 0.0500\text{ mol HCl}$ L.R.

$\frac{0.100\text{ mol }Ba(OH)_2}{L} \times 0.3000L = \cancel{0.0300\text{ mol }Ba(OH)_2}$

$0.0500\,\cancel{\text{mol HCl}} \times \frac{-118\text{ kJ}}{2\,\cancel{\text{mol HCl}}} = -2.95\text{ kJ}$ $\boxed{2.95\text{ kJ released}}$

b. Assuming the temperature of both solutions was initially 25.0°C, and the final mixture has a mass of 400.0 g, calculate the final temperature of the product mixture (assume dilute aqueous solutions have a specific heat capacity of 4.184 J/g·°C).

$Q = mC\Delta T$ $T_f = ?$

$2.95\text{ kJ} = Q$

$Q = ?\text{ J}$

2950 J

$\Delta T = \frac{Q}{mC} = \frac{2950\,\cancel{J}}{(400.0\,\cancel{g})(4.184\,\frac{\cancel{J}}{\cancel{g}°C})}$

$\Delta T = 1.76°C$

$+25.0$

26.76

$\boxed{26.8°C}$

$\Delta T = T_f - T_i$

$1.76°C = T_f - 25.0°C$

 ~~???~~ $C_?H_? + O_2 \rightarrow CO_2 + H_2O$

3. Natural gas (CH_4) as well as octane (C_8H_{18}) are used as fuels for internal combustion engines. Calculate the enthalpy of combustion for each of these fuels in (kJ/mole fuel) and (kJ/gram fuel).

$\Delta H = -890.5\,kJ$

Substance	$\Delta H_f°$ (kJ/mol)
$CH_4(g)$	–74.6
$C_8H_{18}(l)$	–208.5
$CO_2(g)$	–393.5
$H_2O(l)$	–285.8

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

$\frac{1}{2}\left(2C_8H_{18}(g) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)\right) \quad \Delta H = \frac{-11023.4\,J}{2}$

$C_8H_{18}(g) + 12.5O_2 \rightarrow 8CO_2(g) + 9H_2O(g) \quad = -5511.7\,kJ$

$CH_4 \quad \frac{-890.5\,kJ}{1\,mol} \times \frac{1\,mol\,CH_4}{16.042\,g\,CH_4} = -55.51\,\frac{kJ}{g\,CH_4}$

$C_8H_{18} \quad \frac{-5511.7\,kJ}{mol\,C_8H_{18}} \times \frac{1\,mol\,C_8H_{18}}{114.22\,g\,C_8H_{18}} = -48.25\,\frac{kJ}{g\,C_8H_{18}}$

Fuel	$\Delta H_{combustion}$ (kJ/mol)	$\Delta H_{combustion}$ (kJ/gram)
CH_4	–890.5	–55.51
C_8H_{18}	–5511.7	–48.25

4. For the reaction below:

ΔH_f → 0 -520 -350.5 -971 (kJ/mol)

$Zn(s) + 2MnO_2(s) \rightarrow ZnO(s) + Mn_2O_3(s) \quad \Delta H_{rxn} = ?$

a. Using the enthalpy of formation data in the appendix (or on saplinglearning.com), calculate the enthalpy for the reaction.

change

$\Delta H = -350.5 + (-971) - (0 + 2 \cdot (-520))$

$= \boxed{-281.5 \text{ kJ/mol}}$

b. Is the reaction endothermic or exothermic?

c. What mass MnO_2 would be required to produce 50.0 J of heat?

$50.0 \cancel{J} \times \frac{1 \cancel{kJ}}{1000 \cancel{J}} \times \frac{2 \cancel{mol\ MnO_2}}{281.5 \cancel{kJ}} \times \frac{86.94 g\ MnO_2}{1 \cancel{mol\ MnO_2}}$

$= \boxed{0.0309 g\ MnO_2}$