## + WS 7.6 Solutions

1. How much heat (in joules) is required to raise the temp of 205 g of water from 15.2˚C to 16.2˚C?

$q = mc\Delta T$ (205) (4.184) (1.0) = **860 J** Ans _____

2. What mass of 67.5˚C iron must be added to 235 g of 5.00˚C water to make the final temp of both come out to be 15.0˚C?

$mc\Delta T = mc\Delta T$ (m) (0.444) (52.5) = (235) (4.184) (10.0) Ans **422 g**

3. How many kJ are released when 205 L of propane gas ($C_3H_8$) is burned at STP conditions?
(remember, at STP, 1 mole = 22.4 L)

$\Delta H_c = -2219.2$ kJ/mol

$$205\ L \times \frac{1\ mol}{22.4\ L} \times \frac{2219.2\ kJ}{1\ mol} = \mathbf{20300\ kJ}$$

Ans _____

4. Use Hess's Law to calculate ΔH for the following reaction: ($\Delta H_f$ $Al_2O_3(s)$ = -1676 kJ/mol)

**4 Al(s) + 3 $MnO_2$(s) ---> 2 $Al_2O_3$(s) + 3 Mn(s)**

x 2 ← 2 Al + 3/2 $O_2$ ---> $Al_2O_3$ ΔH = -1676 "givens"

x 3, flip ← Mn + $O_2$ ---> $MnO_2$ ΔH = -520.

3 $MnO_2$ ---> 3 Mn + 3 $O_2$ ΔH = +1560

+ 4 Al + 3 $O_2$ ---> 2 $Al_2O_3$ ΔH = -3352

**4 Al + 3 $MnO_2$ ---> 2 $Al_2O_3$ + 3 Mn**

**ΔH** = 1560 + (-3352) = **-1792 kJ/mol**

5. Calculate ΔG for the following set of values: **[ ΔH= -288 kJ/mol, ΔS= 125 J/mol•K, T= 700. K ]**
Would this reaction occur spontaneously? you bet!

change to kg

$\Delta G = \Delta H - T\Delta S$

= (-288) - 700.(0.125)

= **- 376 kJ/mol**

6. Calculate ΔG for the composition of hydrogen sulfide gas @ 298 K:

**$H_2$(g) + 1/2 $S_2$(g) ---> $H_2S$(g)**

look up values for "S" from chart → S = (130.6) 1/2 (228.1) (205.6)

calculate ΔS → ΔS = 205.6 - [ 130.6 + 1/2 (228.1) ]

= - 39.05 kJ/mol = - 0.03905 kJ/mol•K

$\Delta G = \Delta H - T\Delta S$

= -23.9 - [(298)(-0.03905)]

= **-12.3 kJ/mol**

7. Above what temperature (K) would reaction #6 occur spontaneously?

use
$\Delta G = \Delta H - T\Delta S$
set ΔG to 0, and solve for T

0 = ΔH - TΔS
0 = -23.9 - (T)(-0.03905)
23.9 = -T (-0.03905)

T = **612 K**

8. What minimum temperature (˚C) is required for silver chloride (AgCl) to spontaneously decompose?

ΔH = + 127 kJ/mol

AgCl ---> Ag + 1/2 $Cl_2$

ΔS = 1/2 (223) + 42.7 - 96.1
= 58.1 J/mol•K (don't forget to change this to kg)

$\Delta G = \Delta H - T\Delta S$
0 = 127 - T (0.0581)
-127 = -T (0.0581)
T = 2186 - 273 = **1913 °C**

9. 1.05 g of methane ($CH_4$) is heated 175.0 g of water by 82.0˚C. Calculate the experimental Hc value in kj/mol.

step 1: calculate heat (q)
q = mcΔT
q = (175.0)(4.184)(82.0)
q = 60000 J = 60.0 kJ

step 2: calculate moles of $CH_4$

$$1.05\ g\ CH_4 \times \frac{1\ mol}{16\ g} = 0.0656\ mol$$

step 3: Hc (kJ / mol)

$$H_c = \frac{60.0\ kJ}{0.0656\ mol} = \mathbf{915\ kJ/mol}$$