NAME KEY DATE HOUR

Section 16.1 Worksheet

Directions: Solve the following problems. You will neeed the reference charts we have been using throughout the chapter to solve some problems.

1. Calculate the amount of heat produced when 11.94 grams of nitric oxide, NO, burn in air to produce nitrogen dioxide according to the following equation: $2NO + O_2 \rightarrow 2NO_2 + 113.04$ kJ

$$\frac{11.94\,\cancel{g\,NO}}{1} \times \frac{1\,mol\,NO}{30.0\,\cancel{g\,NO}} \times \frac{113.04\,KJ}{2\,\cancel{mol\,NO}} = \boxed{22.49\,KJ}$$

2. Calculate the heat change involved in the reaction $CO_2(g) \rightarrow C(s) + O_2(g)$ using the information below.

Flip ~~$H_2O(l) \rightarrow H_2(g) + 1/2O_2(g)$ ΔH=214.3KJ~~
Flip ~~$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$ ΔH=-667.9KJ~~
Flip ~~$C(s) + 2H_2(g) \rightarrow CH_4(g)$ ΔH=-56.1KJ~~

×2 $\cancel{2H_2(g)} + \cancel{O_2(g)} \rightarrow \cancel{2H_2O(l)}$ $(\Delta H = 214.3\,KJ)2$
$CO_2(g) + \cancel{2H_2O(l)} \rightarrow \cancel{CH_4(g)} + \cancel{2O_2(g)}$ $(\Delta H = +667.9\,KJ$
$\cancel{CH_4(g)} \rightarrow C(s) + \cancel{2H_2(g)}$ $\Delta H = +56.1\,KJ$

$CO_2(g) \rightarrow C(s) + O_2(g)$ $\boxed{\Delta H = +295.4\,KJ}$

ethanol (C_2H_5OH)
$\Delta H_{vap} = 38.6\,KJ/mol$

3. Calculate the amount of released when 16.8 grams of ethanol is condensed at its boiling point.

$q = n\Delta H_{vap}$

$$\frac{16.8\,\cancel{g\,C_2H_5OH}}{1} \times \frac{1\,mol\,C_2H_5OH}{46.07\,\cancel{g\,C_2H_5OH}} = 0.3647\,mol$$

$q = (0.3647\,mol)(38.6\,KJ/mol)$
$\boxed{q = 14.08\,KJ}$

4. Calculate the change in enthalpy for the following reaction given the information below.
$NiO(s) + 2HCl(g) \rightarrow NiCl_2(s) + H_2O(g)$

$\cancel{SOCl_2}(g) + NiO(s) \rightarrow \cancel{SO_2}(g) + NiCl_2(s)$ ΔH= -150kJ
Flip ~~$SOCl_2(g) + H_2O(g) \rightarrow SO_2(g) + 2HCl(g)$ ΔH= 27 kJ~~

$2HCl(g) + \cancel{SO_2(g)} \rightarrow \cancel{SOCl_2(g)} + H_2O(g)$ $\Delta H = +27\,KJ$

$NiO(s) + 2HCl(g) \rightarrow NiCl_2(s) + H_2O(g)$ $\boxed{\Delta H = -123\,KJ}$

5. Calculate the mass of water produced when methane is burned in excess air producing 1546 kJ of heat according to the following equation: $CH_4 + 2O_2 \rightarrow CO_2 + 2\,H_2O + 890.2$ kJ

$$\frac{1546\,\cancel{KJ}}{1} \times \frac{1\,\cancel{mol\,CH_4}}{890.2\,\cancel{KJ}} \times \frac{16.043\,g\,CH_4}{1\,\cancel{mol\,CH_4}} = \boxed{27.86\,g\,CH_4}$$

I don't want methane though

$$\frac{1546\,\cancel{KJ}}{1} \times \frac{2\,\cancel{mol\,H_2O}}{890.2\,\cancel{KJ}} \times \frac{18.015\,g\,H_2O}{1\,\cancel{mol\,H_2O}} = \boxed{62.57\,g\,H_2O}$$

This is the Answer.

6. Determine the energy needed when 55.6 grams of aluminum at 43.2 °C is heated to 78.1 °C.

$Q = m \times \Delta T \times C$

$Q = (55.6\,g)(34.9\,°C)(0.897\,J/g°C) = \boxed{1740.6\,J}$

$\Delta H_{fus} = 5.65\ KJ/mol$

7. If a sample of ammonia, NH_3, required 18.2 kJ of energy to melt it, what is the mass of the sample?

$q = n\Delta H_{vap}$

$\frac{18.2KJ}{5.65KJ/mol} = \frac{n(\cancel{5.65KJ/mol})}{\cancel{5.65KJ/mol}}$

$n = \frac{3.22\ \cancel{mol\ NH_3}}{1} \times \frac{17.01g\ NH_3}{1mol\ \cancel{NH_3}} = \boxed{54.77g\ NH_3}$

8. A student places 42.3 grams of ice at 0.0 °C in an insulated bottle. The student adds 255.8 grams of water at 90.0 °C. Determine the final temperature of the mixture. *Harder

$\boxed{65.9°C}$

ice: $m = 42.3g$, $T_1 = 0°C$, $T_2 = ?$

water: $m = 255.8g$, $T_1 = 90°C$, $T_2 = ?$

$q_{ice} = q_{water}$

two steps

$[q\ n\Delta H_{fus}] + [m \times \Delta T \times C] = m \times \Delta T \times C$

$[2.35mol(6.01KJ/mol)] + [(42.3g)(x)(4.18J/g°C)] = [(255.8g)(90°C - x)(4.18J/g°C)]$

$14{,}115J + 176.8\ J/°C\ x = 1069.2\ J/°C\ (90°C - x)$

$14{,}115J + 176.8\ J/°C\ x = 96232\ J - 1069.2\ J/°C\ x$

x-0°C, water 90°C-x; 0°C, x, 90°C

$42.3g \times \frac{1mol}{18.01g} = 2.35mol$

9. Calculate the amount of energy released when converting 34.0 grams of steam at 100.0 °C to liquid water at 20.0 °C.

$\frac{34.0g\ \cancel{H_2O}}{1} \times \frac{1mol\ H_2O}{18.015g} = 1.888mol$

(1) Steam 100°C → water 100°C

$q = n\Delta H_{vap}$

$q = (1.888mol)(40.67KJ/mol)$

$q = 76.78KJ$

(2) water 100°C → water 20°C

$q = m \times \Delta T \times C$

$q = (34g)(80°C)(4.18J/g°C)$

$q = 11369.6J$

$q = 11.369KJ$

$76.78KJ + 11.369KJ = 88.15KJ$

10. It takes 565 Joules of heat to raise a sample of iron from 21.0 °C to 75.4 °C. What is the mass of the sample?

$q = 565J$

$\Delta T = 54.4°C$

$m = ?$

$C = 0.449\ J/g°C$

$q = m \times \Delta T \times C$

$\frac{565J}{(54.4°C)(.449J/g°C)} = \frac{m(\cancel{54.4°C})(0.449J/g°C)}{(\cancel{54.4°C})(.449J/g°C)}$

$\boxed{m = 23.13g}$

11. A piece of stainless steel weighing 1.55 g absorbs 141 J of heat energy when its temperature increases from 22 °C to 200 °C. What is the specific heat of stainless steel?

$m = 1.55g$

$q = 141J$

$\Delta T = 178°C$

$C = ?$

$\frac{q}{m \times \Delta T} = \frac{m \times \Delta T \times C}{m \times \Delta T}$

$C = \frac{q}{m \times \Delta T}$

$C = \frac{141J}{(1.55g)(178°C)} = \boxed{0.511\ J/g°C}$

12. Calculate the change in enthalpy of $PCl_3\ (l) + Cl_2\ (g) \rightarrow PCl_5\ (s)$ given the following reactions:

FLIP ~~$2P\ (s) + 3Cl_2\ (g) \rightarrow 2PCl_3\ (l)$ $\Delta H = -640kJ$~~

$\cancel{2P\ (s)} + \cancel{5}Cl_2\ (g) \rightarrow 2PCl_5\ (s)$ $\Delta H = -886kJ$

$2PCl_3(l) \rightarrow \cancel{2P(s)} + \cancel{3Cl_2}(g)$ $\Delta H = +640KJ$

$\div 2$ $\cancel{2}PCl_3(l) + \cancel{2}Cl_2(g) \rightarrow \cancel{2}PCl_5$ $\Delta H = \frac{-246KJ}{2}$

$\boxed{\Delta H = -123KJ}$

$\frac{\cancel{1246J/°C}\ x}{\cancel{1246J/°C}} = \frac{82117J}{1246J/°C}$

$\boxed{x = 65.9°C}$