Announcements – Lecture XX I– Monday, June 22nd

1. **Final Lab:** Tuesday, June 23th, ISB 155B
   (Pre-Lab Quiz – TA Evaluation in Class Owls)

2. **Exam III:** Friday, June 26th, In Class
   3 or 4 questions will be taken from Lab Owls:-
   3.4 , 4.2 , 4.5 , 5.5 , 5.6

Quiz 17 Class #: ____ Last Name: _____

A chunk of silver weighing 19.7 grams and originally at 97.48°C is dropped into an insulated cup containing 76.6 grams of water at 23.38°C. Assuming that all of the heat is transferred to the water, the final temperature of the water is: 24.44 °C

Heat Capacity : H_2O = 4.184 J/g°C Ag = 0.237 J/g°C

$q_{H_2O} = m \times C \times \Delta T$

$= 76.6\,(4.184)\,\Delta T$

$= 320.49\,(T_f - 23.38)$

$= 320.49\,T_f - 7493.16$

$q_{Ag} = m \times C \times \Delta T$

$19.7\,(0.237)\,\Delta T$

$4.6689\,(T_f - 97.48)$

$4.6689\,T_f - 455.12$

$\Sigma q_{'s} = 0$

$q_{H_2O} + q_{Ag} = 0$

$320.49\,T_f - 7493.16 + 4.6689\,T_f - 455.12 = 0$

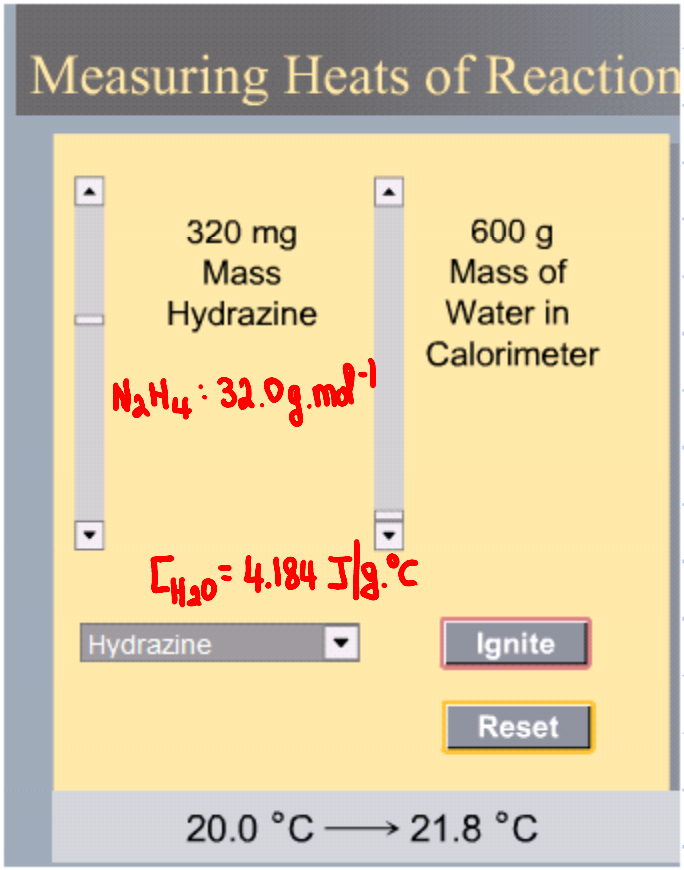
$325.16\,T_f - 7948.28 = 0$

$325.16\,T_f = 7948.28$

$T_f = \frac{7948.28}{325.16} = 24.44°C$

5.4 Enthalpy Changes and Chemical Reactions

C: Determining Enthalpy Change – Calorimetry – 1st Approximation



N_2H_4 : 32.0 g.mol^{-1}

C_{H_2O} = 4.184 J/g.°C

$\Delta T = 21.8 - 20.0 = 1.8°C$

$q_{H_2O} = m \times C \times \Delta T$

$= 600 \times 4.184 \times 1.8 = 4.52 \times 10^3$ J

Σq's = 0 : $q_{RXN} + q_{H_2O} = 0$

$q_{RXN} = -q_{H_2O}$

$= -4.52 \times 10^3$ J

$320\ mg \left| \frac{1\ g}{1000\ mg} \right. = 0.32\ g$

$0.32\ g \left| \frac{1\ mol}{32.0\ g} \right. = 0.01\ mol$

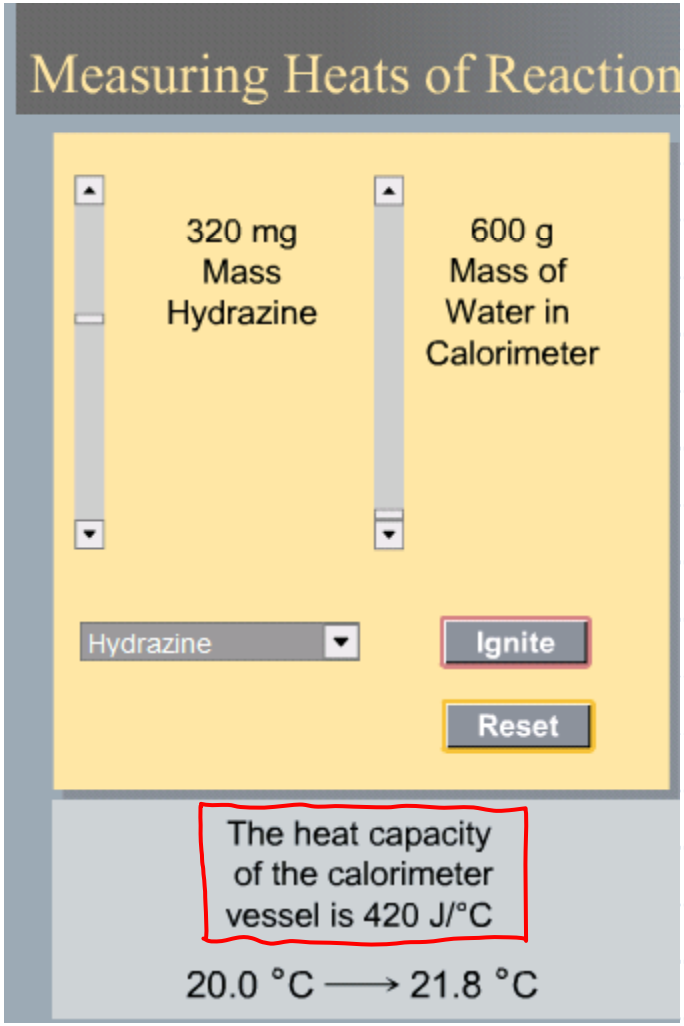
$q_{RXN} = \frac{-4.52 \times 10^3\ J}{0.01\ mol} = -4.52 \times 10^5\ J.mol^{-1}$

OR

$-452\ kJ.mol^{-1}$

5.4 Enthalpy Changes and Chemical Reactions

C: Determining Enthalpy Change – Calorimetry – Modified



q_{H_2O} = 4.52×10^3 J ... see previous slide

q_{cal} = m x C x ΔT

↳ Calorimeter Constant

"Heat capacity of the calorimeter

= 420 J/°C x 1.8°C = 756 J

$\Sigma q_{is} = 0$

$q_{RXN} + q_{H_2O} + q_{cal} = 0$

$q_{RXN} = -(q_{H_2O} + q_{cal})$

$= -(4.52 \times 10^3 + 756)$

$= -5.27 \times 10^3$ J

$q_{RXN} = \frac{-5.27 \times 10^3 \text{ J}}{0.01 \text{ mol}} = -5.27 \times 10^5 \text{ J.mol}^{-1}$

OR

-527 kJ.mol^{-1}

↑ see previous slide

5.4 Enthalpy Changes and Chemical Reactions

C: Determining Enthalpy Change – Calorimetry

A 0.242g sample of napthalene ($C_{10}H_8$) is burned in a bomb calorimeter containing 1025g of water. How much will the temperature rise

Heat capacity of the calorimeter = 802 J/°C
Heat of combustion $C_{10}H_8$ = -5.15x10³ kJ/mol
Heat capacity of water = 4.184 J/g.°C

$C_{10}H_8$: $10(12.01) + 8(1.01) = 128.18\ g.mol^{-1}$

$$0.242\ g \left| \frac{1\ mol}{128.18\ g} \right. = 1.89 \times 10^{-3}\ mol$$

$q_{RXN} = -5.15 \times 10^6\ J.mol^{-1}\ (1.89 \times 10^{-3}\ mol)$
$= -9.72 \times 10^3\ J$

$q_{RXN} + q_{H_2O} + q_{cal} = 0$

$q_{H_2O} + q_{cal} = -q_{RXN} = 9.72 \times 10^3\ J$

$q_{H_2O} = m \times C \times \Delta T$
$= 1025 \times 4.184 \times \Delta T = 4.29 \times 10^3 \Delta T$

$q_{cal} = 802 \Delta T$

$q_{H_2O} + q_{cal} = 9.72 \times 10^3$

$4.29 \times 10^3 \Delta T + 802 \Delta T = 9.72 \times 10^3$

$5.09 \times 10^3 \Delta T = 9.72 \times 10^3$

$$\Delta T = \frac{9.72 \times 10^3}{5.09 \times 10^3} = 1.91\ ^\circ C$$

5.5 Hess's Law

A: Hess's Law

Given the standard enthalpy changes for the following two reactions:

(1) $2\ Pb(s) + O_2(g) \longrightarrow 2\ PbO(s)$ $\quad \Delta H° = -434.6$ kJ

(2) $Pb(s) + Cl_2(g) \longrightarrow PbCl_2(s)$ $\quad \Delta H° = -359.4$ kJ

what is the standard enthalpy change for the reaction:

(3) $2\ PbCl_2(s) + O_2(g) \longrightarrow 2\ PbO(s) + 2\ Cl_2(g)$ $\quad \Delta H° = ?$

2 ... x2 + Reversed $\quad 2\ PbCl_2(s) = 2\ Pb(s) + 2\ Cl_2(g) \quad \Delta H° = -2(-359.4)$

1 ... as is $\quad 2\ Pb(s) + O_2(g) = 2\ PbO(s) \quad \Delta H° = -434.6$

$2\ PbCl_2(s) + O_2(g) = 2\ PbO(s) + 2\ Cl_2(g) \quad \Delta H° = 284.2$ kJ

Hess's Law:

Reverse a reaction ... Reverse the sign of $\Delta H°$.

Multiply reaction by an integer ... Multiply $\Delta H°$ by the integer.

Add two or more reactions ... Add up the $\Delta H°$'s of each.

5.5 Hess's Law

A: Hess's Law

Given the standard enthalpy changes for the following two reactions:

(1) $Ni(s) + Cl_2(g) \longrightarrow NiCl_2(s)$ $\quad \Delta H^\circ = -305.3$ kJ

(2) $Pb(s) + Cl_2(g) \longrightarrow PbCl_2(s)$ $\quad \Delta H^\circ = -359.4$ kJ

what is the standard enthalpy change for the reaction:

(3) $Ni(s) + PbCl_2(s) \longrightarrow NiCl_2(s) + Pb(s)$ $\quad \Delta H^\circ = ?$

1. (1) is
2. Reversed

$Ni(s) + Cl_2(g) = NiCl_2(s)$ $\quad \Delta H^\circ = -305.3$

$PbCl_2(s) = Pb(s) + Cl_2(g)$ $\quad \Delta H^\circ = 359.4$

$Ni(s) + PbCl_2(s) = NiCl_2(s) + Pb(s)$ $\quad \Delta H^\circ = 54.1$ kJ

$\Delta H^\circ_{RXN} = \Delta H^\circ_1 - \Delta H^\circ_2$