



Name: \_\_\_\_\_

**THERMODYNAMICS  
HONORS CHEMISTRY 2005.1.0**

1. How much heat must be added to change the temperature of 250 g of water from 25 °C to 60 °C?

**Solution:**

Using the equation just discussed

$$\text{heat} = \text{s.h.} \cdot m \cdot \Delta T$$

$$\text{heat} = 4.184 \frac{\text{J}}{\text{g} \cdot ^\circ\text{C}} \cdot 250 \text{ g} \cdot (60^\circ\text{C} - 25^\circ\text{C})$$

$$\text{heat} = 36.6 \text{ kJ}$$

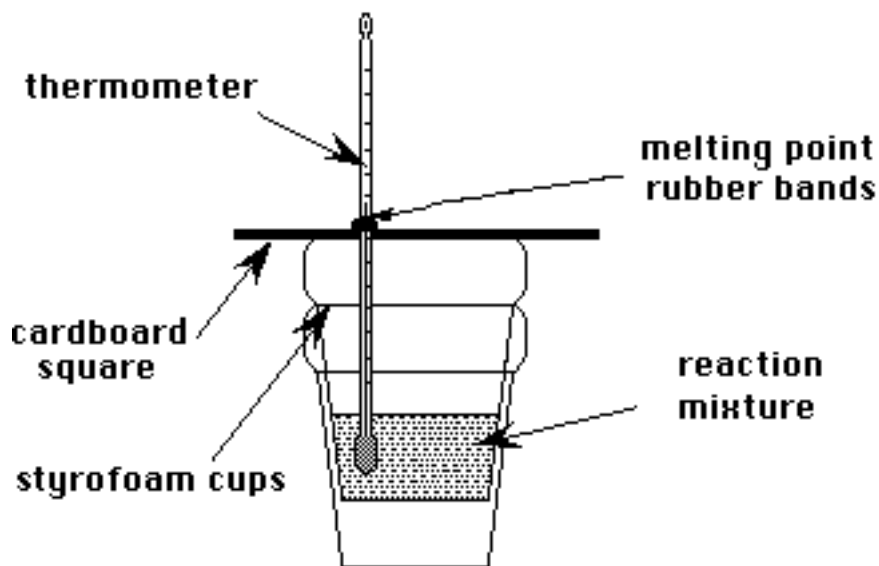
This is the amount of energy required to heat a cup of water for instant coffee.

2. If 2.09 joules are required to change the temperature of 15.0 g of mercury by 1.00 °C, calculate the specific heat of mercury.

**Solution:**

$$\text{Specific heat} = \frac{2.09 \text{ joules}}{15.0 \text{ g} \cdot 1.00^\circ\text{C}} = 0.139 \frac{\text{J}}{\text{g} \cdot ^\circ\text{C}}$$

3. Label the important components of the coffee-cup calorimeter shown below. What property of a chemical reaction does a coffee-cup calorimeter measure? What type of reactions are studied? **See Appendix III for recommended demonstration, video, or computer resources.**



The heat given off or absorbed when using a coffee-cup calorimeter is equal to the change in enthalpy at constant pressure for the chemical reaction studied.

A 'coffee-cup' calorimeter is used to measure the enthalpy for a chemical reaction which occurs in water. Generally reactions in solution are studied. Examples include neutralization reactions and heats of solution.

4. Write the mathematical equation that relates the heat released by the chemical reaction to the heat absorbed by the water and the coffee-cup calorimeter.

$$q_{\text{rxn}} = -(q_{\text{calorimeter}} + q_{\text{water}})$$

$$q_{\text{rxn}} = -((C_{\text{calorimeter}} \cdot \Delta T) + (\text{S.H.}_{\text{water}} \cdot \text{mass}_{\text{water}} \cdot \Delta T))$$

$C_{\text{calorimeter}}$  is the heat capacity of the calorimeter

$\Delta T$  is the change in temperature as a result of the chemical reaction

$\text{S.H.}_{\text{water}}$  is the specific heat of water (Note: Even though a solution is formed, one generally uses the specific heat of water as a first approximation.)

$\text{mass}_{\text{water}}$  is the mass of the water in the calorimeter

$$q_{\text{rxn}} = \Delta H_{\text{rxn}}$$

5. When 1.095 g of NaOH is dissolved in 150.0 g of water initially at 23.50 °C in a coffee-cup calorimeter, the final temperature is found to be 25.32 °C. Calculate the heat liberated when NaOH dissolves in water. (Assume the specific heat of the solution is the same as that of water and no heat is absorbed by the calorimeter.)

$$\text{heat} = \text{s.h.} \cdot m \cdot \Delta T$$

$$\text{heat} = 4.184 \frac{\text{J}}{\text{g} \cdot ^\circ\text{C}} (151.1 \text{ g})(25.32 - 23.50) = 1,151 \text{ J liberated}$$

6. A 50.00 g sample of 0.200 M NaBr(aq) at 23.65 °C is added to a coffee-cup calorimeter containing 50.00 g of 0.200 M AgNO<sub>3</sub>(aq) at 23.65 °C. If the heat capacity of the calorimeter is 65.0  $\frac{\text{J}}{^\circ\text{C}}$  and the specific heat of the solution is 4.20  $\frac{\text{J}}{\text{g} \cdot ^\circ\text{C}}$  and the final temperature of the solution in the calorimeter is 25.40 °C, calculate the heat released in the reaction.

$$-(q_{\text{solution}} + q_{\text{calorimeter}}) = q_{\text{reaction}}$$

$$q_{\text{solution}} = (\text{specific heat})_{\text{solution}} \cdot \text{mass}_{\text{solution}} \cdot \Delta T_{\text{solution}}$$

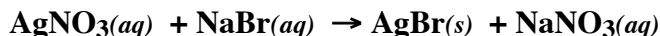
$$q_{\text{solution}} = 4.20 \frac{\text{J}}{\text{g} \cdot ^\circ\text{C}} \cdot 100.0 \text{ g} \cdot (25.40 ^\circ\text{C} - 23.65 ^\circ\text{C}) = 7.35 \times 10^2 \text{ J}$$

$$q_{\text{calorimeter}} = (\text{heat capacity})_{\text{calorimeter}} \cdot \Delta T_{\text{calorimeter}}$$

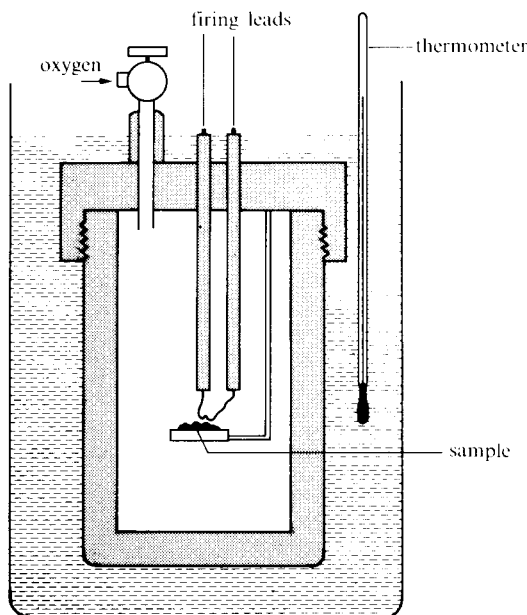
$$q_{\text{calorimeter}} = 65.0 \frac{\text{J}}{^\circ\text{C}} \cdot (25.40 ^\circ\text{C} - 23.65 ^\circ\text{C}) = 1.14 \times 10^2 \text{ J}$$

$$q_{\text{reaction}} = -(7.35 \times 10^2 \text{ J} + 1.14 \times 10^2 \text{ J}) = -8.49 \times 10^2 \text{ J}$$

This is the heat released when the reaction occurred. The reaction is



7. Label the important components of a bomb calorimeter shown below. What property of a chemical reaction does a bomb calorimeter measure? What type of reaction is studied? See Appendix III for recommended demonstration, video, or computer resources.



A bomb calorimeter is used to determine  $\Delta E$  for a chemical reaction. It is generally used to measure the internal energy of a combustion reaction. A measured amount of a sample to be combusted is placed on a sample holder in the 'bomb'. The bomb is sealed and dioxygen is pumped into the bomb. The bomb is placed in a large reservoir of water into which an accurate thermometer is placed. Electrical leads are attached to a battery and are used to spark the dioxygen. The heat liberated from the reaction flows into the water and the temperature change is measured.

8. Write the mathematical equation that relates the heat released by the chemical reaction to the heat absorbed by the water and the bomb calorimeter.

The heat given off or absorbed when using a bomb calorimeter is equal to the change in internal energy ( $\Delta E$ ) for the chemical reaction studied.

$$q_{\text{rxn}} = -(q_{\text{calorimeter}} + q_{\text{water}})$$

$$q_{\text{rxn}} = -((C_{\text{calorimeter}} \cdot \Delta T) + (S.H._{\text{water}} \cdot \text{mass}_{\text{water}} \cdot \Delta T))$$

$C_{\text{calorimeter}}$  is the heat capacity of the calorimeter

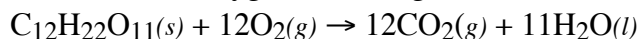
$\Delta T$  is the change in temperature as a result of the chemical reaction

$S.H._{\text{water}}$  is the specific heat of water

$\text{mass}_{\text{water}}$  is the mass of the water in the calorimeter

$$q_{\text{rxn}} = \Delta E_{\text{rxn}}$$

9. Sucrose,  $C_{12}H_{22}O_{11}$  reacts with oxygen according to the reaction,



Calculate the heat produced per mole of sucrose when 2.75 g of  $C_{12}H_{22}O_{11}$  are combusted with excess oxygen in a bomb calorimeter containing 4.80 kg of water. The temperature change measured is  $2.01\text{ }^\circ\text{C}$ . The heat capacity of the calorimeter is  $2540\frac{\text{J}}{^\circ\text{C}}$ .

**Solution:**

**The heat given off by the reaction is equal to the heat absorbed by the water in the calorimeter and the calorimeter.**

$$q_{\text{reaction}} = -(q_{\text{water}} + q_{\text{calorimeter}})$$

**The amount of heat absorbed by the water is given by**

$$q_{\text{water}} = (\text{specific heat})_{\text{water}} \cdot \text{mass}_{\text{water}} \cdot \Delta T_{\text{water}}$$

$$q_{\text{water}} = 4.184\frac{\text{J}}{\text{g}\cdot^\circ\text{C}} \cdot 4800\text{ g} \cdot 2.01\text{ }^\circ\text{C}$$

$$q_{\text{water}} = 40.4\text{ kJ}$$

**The amount of heat absorbed by the calorimeter is given by**

$$q_{\text{calorimeter}} = (\text{heat capacity})_{\text{calorimeter}} \cdot \Delta T_{\text{calorimeter}}$$

$$q_{\text{calorimeter}} = 2540\frac{\text{J}}{^\circ\text{C}} \cdot 2.01\text{ }^\circ\text{C}$$

$$q_{\text{calorimeter}} = 5.10\text{ kJ}$$

$$q_{\text{reaction}} = -(q_{\text{water}} + q_{\text{calorimeter}})$$

$$q_{\text{reaction}} = -(40.4\text{ kJ} + 5.10\text{ kJ})$$

$$q_{\text{reaction}} = -45.5\text{ kJ}$$

**The amount of heat produced per mole of sucrose is,**

$$\left(\frac{45.5\text{ kJ}}{2.75\text{ g}}\right)\left(\frac{342.0\text{ g}}{1\text{ mol } C_{12}H_{22}O_{11}}\right) = 5.66 \times 10^3\frac{\text{kJ}}{\text{mol}}\text{ released}$$

$$q_{\text{water}} = \Delta E = -5.66 \times 10^3\text{ kJ}$$

10. A 28.4 g sample of an unknown metal was heated to 110.0 °C and plunged into a 100 g sample of water initially at a temperature of 24.60 °C. The final temperature of the mixture was 25.34 °C. Calculate the specific heat of the metal. Identify the metal.

$$q = m \cdot s \cdot h \cdot \Delta T$$

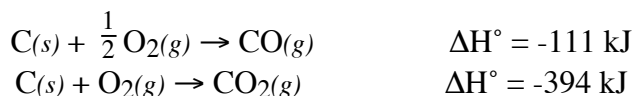
$$q_{\text{gain by H}_2\text{O}} = -q_{\text{lost by metal}}$$

$$\begin{aligned} \text{S.H. metal} &= \frac{q_{\text{gained by H}_2\text{O}}}{-M_{\text{metal}} \Delta T_{\text{metal}}} \\ &= - \left( \frac{(100.0 \text{ g})(4.184 \frac{\text{J}}{\text{g} \cdot \text{C}})(25.34 \text{ C} - 24.60 \text{ C})}{28.4 \text{ g}(25.34 \text{ C} - 110.0 \text{ C})} \right) \\ &= - \left( \frac{309.616 \text{ J}}{-2404.344 \text{ g} \cdot \text{C}} \right) \\ &= 0.129 \frac{\text{J}}{\text{g} \cdot \text{C}} ; \text{ metal could be Au - had same specific heat} \end{aligned}$$

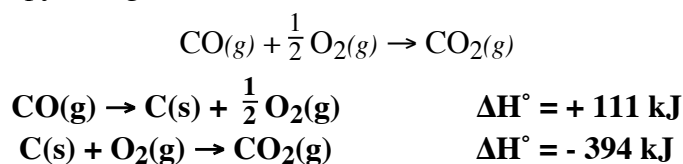
11. Define *Hess' Law*.

**Hess's Law - in going from a particular set of reactants to a particular set of products, the enthalpy change is the same whether the reaction takes place in one step or in a series of steps.  $\Delta H$  for the overall reaction will be equal to the sum of the enthalpy changes of each step.**

12. Given the enthalpy change for the two reactions below,



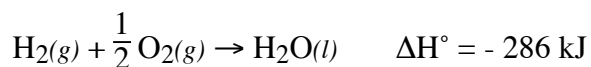
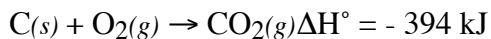
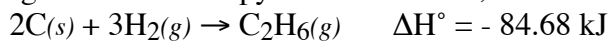
Calculate the enthalpy change for the reaction,



**Adding the two equations**

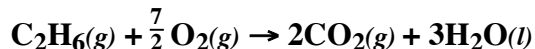


13. Using the following standard enthalpy of reaction data,

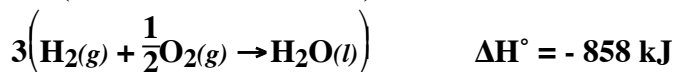
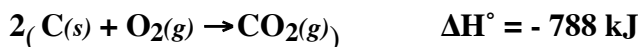
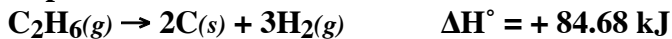


Calculate the heat of reaction for the combustion of 1 mol of ethane ( $\text{C}_2\text{H}_6$ ).

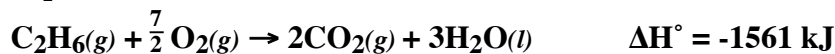
**The equation which describes the combustion of ethane is,**



**Rearranging the equations provided**



**Adding the equations**

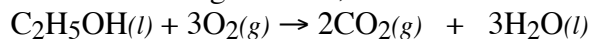


14. Define the term *heat of reaction* and write the mathematical equation used to calculate the heat of a chemical reaction.

**The heat (enthalpy) of reaction can be determined by adding together the enthalpies of formation associated with a series of formation reactions which can generate the desired reaction.**

$$\Delta H_{\text{rxn}}^\circ = \sum \Delta H_f^\circ (\text{products}) - \sum \Delta H_f^\circ (\text{reactants})$$

15. Calculate the  $\Delta H^\circ$  for the following reaction,



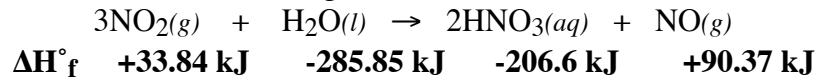
$$\Delta H_f^\circ \quad -277.7 \text{ kJ} \quad 0 \quad -393.5 \text{ kJ} \quad -285.85 \text{ kJ}$$

$$\Delta H_{\text{rxn}}^\circ = \sum \Delta H_f^\circ (\text{products}) - \sum \Delta H_f^\circ (\text{reactants})$$

$$\Delta H_{\text{rxn}}^\circ = 3 \text{ mol} \left( -285.85 \frac{\text{kJ}}{\text{mol}} \right) + 2 \text{ mol} \left( -393.5 \frac{\text{kJ}}{\text{mol}} \right) - \left( 1 \text{ mol} \left( -277.7 \frac{\text{kJ}}{\text{mol}} \right) \right)$$

$$\Delta H_{\text{rxn}}^\circ = -1366.8 \text{ kJ}$$

16. Calculate the  $\Delta H^\circ$  for the following reaction,

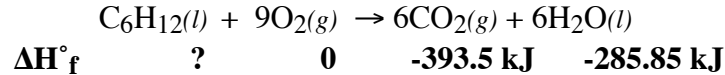


$$\Delta H^\circ_{\text{rxn}} = \Sigma \Delta H^\circ_f (\text{products}) - \Sigma \Delta H^\circ_f (\text{reactants})$$

$$\Delta H^\circ_{\text{rxn}} = \left( 2 \text{ mol} \left( -206.6 \frac{\text{kJ}}{\text{mol}} \right) + 1 \text{ mol} \left( +90.37 \frac{\text{kJ}}{\text{mol}} \right) \right) - \left( 3 \text{ mol} \left( +33.84 \frac{\text{kJ}}{\text{mol}} \right) + 1 \text{ mol} \left( -285.85 \frac{\text{kJ}}{\text{mol}} \right) \right)$$

$$\Delta H^\circ_{\text{rxn}} = -138.5 \text{ kJ}$$

17. The standard enthalpy of combustion to  $\text{CO}_2(g)$  and  $\text{H}_2\text{O}(l)$  at  $25^\circ\text{C}$  of cyclohexane,  $\text{C}_6\text{H}_{12}(l)$ , is  $-3924 \frac{\text{kJ}}{\text{mol}}$ . Calculate the standard heat of formation,  $\Delta H^\circ_f$ , of cyclohexane.



$$\Delta H^\circ_{\text{rxn}} = \Sigma \Delta H^\circ_f (\text{products}) - \Sigma \Delta H^\circ_f (\text{reactants})$$

$$\Delta H^\circ_{\text{rxn}} = 6\Delta H^\circ_f (\text{CO}_2) + 6\Delta H^\circ_f (\text{H}_2\text{O}) - \Delta H^\circ_f (\text{C}_6\text{H}_{12})$$

$$\Delta H^\circ_f (\text{C}_6\text{H}_{12}) = 6\Delta H^\circ_f (\text{CO}_2) + 6\Delta H^\circ_f (\text{H}_2\text{O}) - \Delta H^\circ_{\text{rxn}}$$

$$\Delta H^\circ_f (\text{C}_6\text{H}_{12}) = \left( 6 \text{ mol} \left( -393.5 \frac{\text{kJ}}{\text{mol}} \right) + 6 \text{ mol} \left( -285.85 \frac{\text{kJ}}{\text{mol}} \right) \right) - (-3924 \text{ kJ})$$

$$\Delta H^\circ_f (\text{C}_6\text{H}_{12}) = -152.1 \frac{\text{kJ}}{\text{mol}}$$