

Chemistry 103, Dr. Hamers  
Help Sheet #7, Chapter 5 Concepts  
Tony Jacob

**Nuggets:** terms; heat capacity;  $q = Cm\Delta T$ ; heat transfer (heat lost + heat gained = 0);  $\Delta H_{rxn}$ ; Hess's law;  $\Delta H_f$ ; phase changes ( $\Delta H_{fusion}$  and  $\Delta H_{vaporization}$ ); calorimetry; heat transfer + phase change

### TERMS

Energy: Capacity to do work or to produce heat

Potential energy: energy from position or composition (height above ground, etc.)

Kinetic energy: energy due to motion ( $KE = \frac{1}{2}mv^2$ ) where  $v$  = velocity (Note: also ( $KE = \frac{3}{2}RT$ ))

State function: a property that is independent of the pathway chosen (capitalized/uppercase variables are state functions; lowercase variable are not state functions)

System: the part of the universe that is under observation

Surroundings: everything that is not the system

**Exothermic:** energy is given off (i.e., heat is given off, the solution heats up)

**Endothermic:** energy is absorbed (i.e., heat is absorbed, the solution cools down)

First Law of Thermodynamics: Energy is conserved.

Enthalpy:  $H$  is enthalpy with  $H = U + PV$  where  $U$  is internal energy and  $PV$  is pressure times volume (for most purposes, enthalpy is treated as the heat of the system, this occurs when the pressure of the system is held constant)

**Work and the 1st Law of Thermodynamics:** energy is conserved:  $\Delta U = q + w$  ( $q$  = heat;  $w$  = work)

Work is defined as PV-work which is:  $w = -P\Delta V$  and

Enthalpy is defined as:  $\Delta H = \Delta U + P\Delta V$  and substituting  $\Delta E = q + w$  and at constant  $P$  this becomes:

$$\Delta H = q_p + w + (-w); \quad \Delta H = q_p$$

Back to  $\Delta U = q + w$  and substituting  $w = -P\Delta V$  yields:  $\Delta U = q - P\Delta V$  and at constant  $V$ :  $\Delta U = q_v$

Again using  $\Delta U = q - P\Delta V$  at constant  $P$  yields:  $\Delta U = q_p - P\Delta V$  then substituting:

$$\Delta U = q_v \text{ and } w = -P\Delta V \text{ yields: } q_v = q_p + w \text{ and rearranging yields: } q_p = q_v - w$$

work done by the system  $< 0$  (-); work done on the system  $> 0$  (+)

$$\Delta U = q + w ; \Delta H = q_p ; \Delta U = q_v ; q_p = q_v - w$$

### HEAT CAPACITY

Heat capacity: denoted by a  $C$  with units of  $J/^\circ C$  where  $C = \frac{\text{heat absorbed}}{\text{increase in } T}$

Specific heat capacity: with units of  $J/g^\circ C$ ; the energy required to raise 1 g of material  $1^\circ C$

$$\text{Specific heat capacity} = \frac{\text{heat absorbed}}{\text{increase in } T \times \text{grams of material}}$$

$$q = Cm\Delta T \text{ where } \Delta T = T_f - T_i; C = \text{specific heat capacity (J/g}^\circ\text{C)}, m = \text{mass (g)}; q = \text{heat (J)}$$

**HEAT TRANSFER:** Heat lost + Heat gained = 0

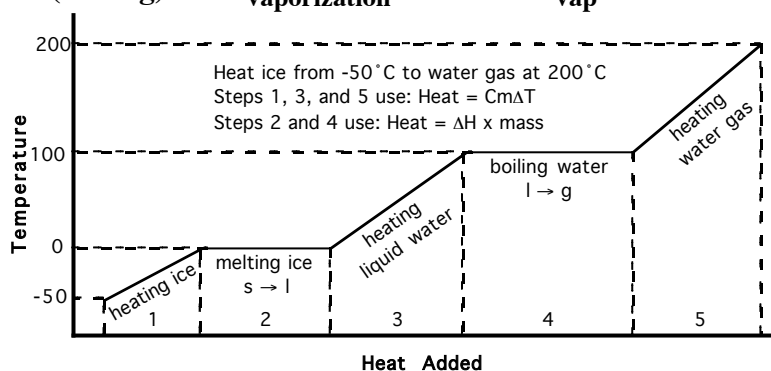
$$C_1 \times \text{mass}_1 \times (T_f - T_{i1}) + C_2 \times \text{mass}_2 \times (T_f - T_{i2}) = 0; (T_{f1} = T_{f2} \text{ in thermal equilibrium})$$

**PHASE CHANGE:** There is no temperature change during a phase change!

Terms: freezing, melting, boiling, condensing

Solid  $\leftrightarrow$  Liquid: **Heat (melting) =  $\Delta H_{\text{fusion}} \times \text{mol} = \Delta H_{\text{fus}} \times \text{mol}$**

Liquid  $\leftrightarrow$  Gas: **Heat (boiling) =  $\Delta H_{\text{vaporization}} \times \text{mol} = \Delta H_{\text{vap}} \times \text{mol}$**



## CALORIMETRY

Typical steps:

Given reaction with a known  $\Delta H_{\text{rxn}}$  find  $T_f$ :

1. Find limiting reactant and determine how many moles of chemicals reacted
2. Use proportionality to find how much heat evolved/absorbed when reaction happened
3.  $q_{\text{sys}} = -q_{\text{surr}}$
4. Heat =  $C_{\text{water}}m(T_f - T_i)$  – find  $T_f$

Given reaction with a known  $T_f$  and find  $\Delta H_{\text{rxn}}$ :

1. Heat =  $C_{\text{water}}m(T_f - T_i)$  – find Heat
2.  $q_{\text{sys}} = -q_{\text{surr}}$
3. Find limiting reactant and determine how many moles of chemicals reacted
4. Use proportionality to find  $\Delta H_{\text{rxn}}$

## ENTHALPY

Change in Enthalpy:  $\Delta H$ ;  $\Delta H = H_f - H_i$  with i = initial conditions and f = final conditions

The change in enthalpy for a reaction,  $\Delta H_{\text{rxn}}$  is expressed as:

$$\Delta H_{\text{rxn}} = H_{\text{products}} - H_{\text{reactants}}$$

**HESS'S LAW:** If two or more reactions are added, the  $\Delta H$  of the net reaction is the sum of the individual  $\Delta H$  for each reaction added;

1. Reverse reaction: change the sign of  $\Delta H$ ;  $\Delta H_{\text{forward}} = -\Delta H_{\text{reverse}}$
2. Multiply the coefficients of a rxn: multiply  $\Delta H_{\text{rxn}}$  by the same factor
3.  $\Delta H_{\text{net rxn}} = \Delta H_{\text{rxn1}} + \Delta H_{\text{rxn2}} + \dots$  (Hess's Law)

## $\Delta H_f^\circ$

$\Delta H_f^\circ$  is the enthalpy of formation for 1 mole of material at standard conditions; The definition (which you should memorize) for  $\Delta H_f^\circ$  is: The heat is required to form 1 mole of the a substance from its elements in their elemental or natural state under standard conditions (25 °C, 1 atm, 1M)

The ° on  $\Delta H$  ( $\Delta H^\circ$ ) means standard conditions: 25 °C; 1 atm pressure, and all solutions at 1M

**$\Delta H_f^\circ$  FOR ALL ELEMENTS IN THEIR ELEMENTAL FORM IS ZERO.**

Use  $\Delta H_f^\circ$  to calculate the  $\Delta H_{\text{rxn}}^\circ$  for a reaction:  $\Delta H_{\text{rxn}}^\circ = \sum n_p \Delta H_f^\circ(\text{products}) - \sum n_r \Delta H_f^\circ(\text{reactants})$

where  $\Sigma$  (sigma) means the sum of,  $n_p$  means the number of moles of each product, and  $n_r$  means the number of moles of each reactant

*Diatomic elements:*  $\text{F}_2, \text{Cl}_2, \text{Br}_2, \text{I}_2, \text{H}_2, \text{N}_2, \text{O}_2$

*Liquid elements:*  $\text{Br}_2, \text{Hg}$ ; *Gaseous elements:*  $\text{He, Ne, Ar, Kr, Rn, F}_2, \text{Cl}_2, \text{H}_2, \text{N}_2, \text{O}_2$

*Rest are solids;* Standard state:  $\text{C} = \text{graphite}$

## PHASE CHANGE + HEAT TRANSFER

Examples:

- ice at  $-20^\circ\text{C}$  heats up, then melts in hot water:  $q_{\text{heat the ice}} + q_{\text{melt ice}} + q_{\text{heat water from ice}} + q_{\text{cool down water}} = 0$
- cold metal placed in warm water that freezes:  $q_{\text{heat the metal}} + q_{\text{cool down water}} + q_{\text{freeze the water}} = 0$
- hot metal into cool water which all boils:  $q_{\text{cool the metal}} + q_{\text{heat the water}} + q_{\text{boil the water}} + q_{\text{heat the water gas}} = 0$

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1. An iron skillet weighing 1510g is heated on a stove to  $178^\circ\text{C}$ . Suppose the skillet is cooled to room temperature,  $21^\circ\text{C}$ . How much heat in joules was given off if the specific heat of iron is  $0.450\text{J/g } ^\circ\text{C}$ ?

2. Exactly 313.5J will raise the temperature of 10.0g of a metal from  $25.0^\circ\text{C}$  to  $60.0^\circ\text{C}$ . What is the specific heat of the metal?

3. Suppose you wish to heat water to make coffee. How much heat must be used to raise the temperature of 0.180 kg of water (about 1 cup) from  $15.0^\circ\text{C}$  (cold tap water temperature) to  $96.0^\circ\text{C}$  (brewing temperature)? Assume the specific heat of water is  $4.184\text{ J/g } ^\circ\text{C}$ .

4. A 5.00 grams piece of Cu with a specific heat capacity of  $0.385\text{J/g } ^\circ\text{C}$  is at  $100.0^\circ\text{C}$ . It is dropped into 10.0g of water at  $20.0^\circ\text{C}$  and it has a specific heat capacity of  $4.184\text{J/g } ^\circ\text{C}$ , what is the final temperature of the water?

5. Describe the following processes as exothermic or endothermic.

- $\text{O}_2(\text{g}) \rightarrow \text{O}_2(\text{l})$
- $\text{O}_2(\text{g}) \rightarrow 2\text{O}(\text{g})$
- $\text{CO}_2(\text{s}) \rightarrow \text{CO}_2(\text{g})$
- $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g})$

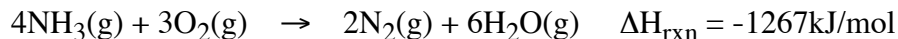
6. How much heat is required to convert 10.0g of ice at  $-25.0^\circ\text{C}$  into water at  $50.0^\circ\text{C}$ ? The specific heat capacity of ice is  $2.10\text{J/g } ^\circ\text{C}$ , the specific heat capacity of water is  $4.184\text{J/g } ^\circ\text{C}$ , and the heat of fusion for ice is  $333\text{J/g}$ .

7. How much heat is require to raise 50ml of methanol,  $\text{CH}_3\text{OH}$ , ( $D = 0.80\text{g/ml}$ ) from  $-25^\circ\text{C}$  to  $100.^\circ\text{C}$  given the following data?  $m_p = -97^\circ\text{C}$ ;  $b_p = 65^\circ\text{C}$ ; specific heat capacity of  $\text{CH}_3\text{OH}(\text{s}) = 1.5\text{J/g } ^\circ\text{C}$  (est.); specific heat capacity of  $\text{CH}_3\text{OH}(\text{l}) = 2.48\text{J/g } ^\circ\text{C}$ ; specific heat capacity of  $\text{CH}_3\text{OH}(\text{g}) = 1.63\text{J/g } ^\circ\text{C}$ ; heat of fusion of  $\text{CH}_3\text{OH} = 99.2\text{J/g}$ ; heat of vaporization of  $\text{CH}_3\text{OH} = 1101\text{J/g}$

8. A reaction between NaOH and HCl occurs by adding 100.ml of a 0.500M NaOH solution to a 250.ml 1.00M HCl solution. Both solutions are at  $15.0^\circ\text{C}$ . The specific heat capacity of water is  $4.18\text{J/g } ^\circ\text{C}$ . Assume the density of the solution is  $1.00\text{g/ml}$ . The  $\Delta H_{\text{rxn}}$  for the reaction between NaOH and HCl is  $-56.0\text{kJ/mol}$ . What is the final temperature of the water after the reaction is complete?  $\text{HCl} + \text{NaOH} \rightarrow \text{H}_2\text{O} + \text{NaCl}$

9. What is the final temperature of a solution if 12.5g  $\text{NH}_4\text{NO}_3(\text{s})$  is dissolved in an insulated coffee cup containing 150.g of water at  $50.0^\circ\text{C}$ ? The heat of dissolving  $\text{NH}_4\text{NO}_3(\text{s})$  is  $26.4\text{kJ/mol}$ , and the heat capacity of the solution is  $4.184\text{J/g}^\circ\text{C}$ .  $\text{NH}_4\text{NO}_3(\text{s}) \rightarrow \text{NH}_4\text{NO}_3(\text{aq})$

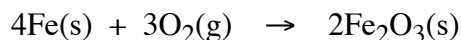
10. Ammonia burns in the presence of a copper catalyst to form nitrogen gas.



a. How much heat would be given off if 1.000 mol of ammonia was consumed?

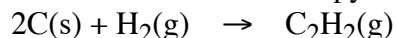
b. How much heat would be given off if 1.000 gram of ammonia was consumed?

11. The overall reaction for the corrosion of iron by oxygen (rusting) is

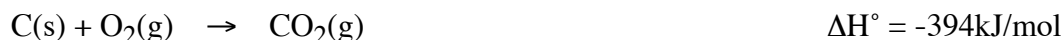
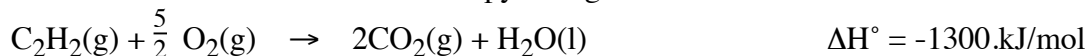


When 10.0g of Fe is reacted, 73.9kJ are released. Calculate the  $\Delta H_{\text{rxn}}$ .

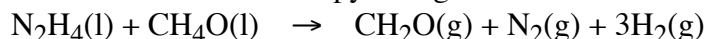
12. Calculate the standard enthalpy change  $\Delta H^\circ$  for



given the individual reaction standard enthalpy changes



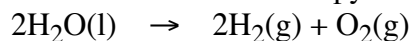
13. Calculate the standard enthalpy change  $\Delta H^\circ$  for



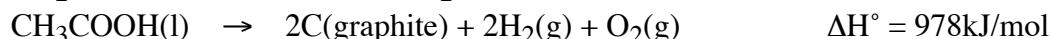
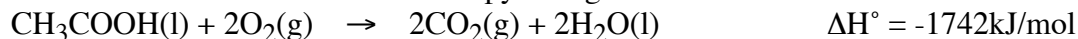
given the individual reaction standard enthalpy changes



14. Calculate the standard enthalpy change  $\Delta H^\circ$  for



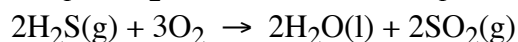
given the individual reaction standard enthalpy changes



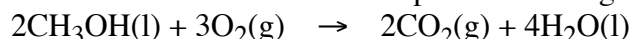
15. Write the  $\Delta H_f$  reaction that corresponds with each chemical listed below.

a.  $\text{CH}_3\text{OH}(\text{l})$  b.  $\text{NH}_4\text{OH}(\text{s})$  c.  $\text{Cu}(\text{NO}_3)_2(\text{s})$

16. Calculate the standard enthalpy change for the reaction below using the standard enthalpies of formation given.  $\Delta H^\circ_f$  for  $\text{H}_2\text{S} = -20.2\text{kJ/mol}$ ,  $\Delta H^\circ_f$  for  $\text{H}_2\text{O} = -285.8\text{kJ/mol}$ ,  $\Delta H^\circ_f$  for  $\text{SO}_2 = -296.8\text{kJ/mol}$



17. The combustion of methanol takes place according to the reaction:



Compute  $\Delta H_{\text{rxn}}^\circ$  for the combustion of **one mole** of methanol given the following standard heats of formation:

$\Delta H^\circ_f$  for  $\text{CH}_3\text{OH} = -238.5\text{kJ/mol}$ ;  $\Delta H^\circ_f$  for  $\text{CO}_2 = -393.5\text{kJ/mol}$ ;  $\Delta H^\circ_f$  for  $\text{H}_2\text{O} = -285.6\text{kJ/mol}$

18. If 50.0g of metal ( $C = 4.00\text{J/g}^\circ\text{C}$ ) initially at  $-100.^\circ\text{C}$  is placed in a 50.0g sample of water at  $75.0^\circ\text{C}$  ( $C = 4.18\text{J/g}^\circ\text{C}$ ), how many grams of water, if any, freezes? The heat of fusion for water is  $333\text{J/g}$ .
19. How much heat is required to take 5.00g of ice at  $-20.0^\circ\text{C}$  to water gas at  $130.^\circ\text{C}$ ? The specific heat of ice is  $2.03\text{J/g}^\circ\text{C}$ , the specific heat of water is  $4.184\text{J/g}^\circ\text{C}$ , the specific heat of  $\text{H}_2\text{O}(\text{g})$  is  $2.00\text{J/g}^\circ\text{C}$ , the heat of fusion for ice is  $333\text{J/g}$ , and the heat of vaporization of water is  $2256\text{J/g}$ .
20. If a beaker contains 5.00g of ice at  $0.00^\circ\text{C}$  and 10.0g of water also at  $0.00^\circ\text{C}$  and 5000.J of heat are added to the beaker, what is the final temperature of the mixture? The specific heat capacity of water is  $4.184\text{J/g}^\circ\text{C}$  and the enthalpy of fusion is  $6.02\text{kJ/mol}$ .
21. a. If a beaker contains 5.00g of ice at  $-10.0^\circ\text{C}$  and 15.0 g of water at  $10.0^\circ\text{C}$ , what is the final temperature of the mixture? The specific heat of ice is  $2.03\text{J/g}^\circ\text{C}$ , the specific heat of water is  $4.184\text{J/g}^\circ\text{C}$ , the specific heat of  $\text{H}_2\text{O}(\text{g})$  is  $2.00\text{J/g}^\circ\text{C}$ , the heat of fusion for ice is  $333\text{J/g}$ , and the heat of vaporization of water is  $2256\text{J/g}$ .  
b. How much if any of the ice melt?

### ANSWERS.

1.  $-1.07 \times 10^5\text{J}$  released    2.  $0.896\text{J/g}^\circ\text{C}$     3.  $6.10 \times 10^4\text{J}$     4.  $23.5^\circ\text{C}$
5. a. endothermic    b. endothermic    c. endothermic    d. exothermic    6. 5960J
7. 55000J    {heat =  $q_{\text{heat liquid}} + q_{\text{boil liquid}} + q_{\text{heat gas}} = (2.48)(50 \times 0.8)(65 - (-25)) + (1101)(50 \times 0.8) + (1.63)(50 \times 0.8)(100 - 65) = 55250\text{J}$ }
8.  $16.9^\circ\text{C}$  (Note: if you got  $1.9^\circ\text{C}$ , that is how much the temperature of the water increased.)
9.  $43.4^\circ\text{C}$     { $1\text{mol NH}_4\text{NO}_3/26.4\text{kJ} = 12.5\text{g NH}_4\text{NO}_3/\text{x}$ ;  $1\text{mol NH}_4\text{NO}_3/26.4\text{kJ} = 0.1563\text{mol NH}_4\text{NO}_3/\text{x}$ ;  $\text{x} = 4.125\text{kJ}$  was absorbed by the  $\text{NH}_4\text{NO}_3$  as it dissolved;  $q_{\text{sys}} = -q_{\text{surr}}$ ;  $q_{\text{surr}} = -4.125\text{kJ} = -4125\text{J}$ ;  $q_{\text{soln}} = C_m\Delta T$ ;  $-4125 = (4.184)(150)(T_f - 50)$ ;  $T_f = 43.4^\circ\text{C}$ }
10. a.  $-316.8\text{kJ}$     b.  $-18.60\text{kJ}$
11.  $-1650\text{kJ/mol}$     12.  $226\text{kJ/mol}$     13.  $-18.5\text{kJ/mol}$  ( $-19\text{kJ/mol}$  with correct sig figs)    14.  $1140\text{kJ/mol}$
15. a.  $\text{C}(\text{graphite}) + 2\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{CH}_3\text{OH}(\text{l})$     b.  $\frac{1}{2}\text{N}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) + \frac{5}{2}\text{H}_2(\text{g}) \rightarrow \text{NH}_4\text{OH}(\text{s})$   
c.  $\text{Cu}(\text{s}) + \text{N}_2(\text{g}) + 3\text{O}_2(\text{g}) \rightarrow \text{Cu}(\text{NO}_3)_2(\text{s})$
16.  $-1125\text{kJ/mol}$     17.  $-726.2\text{kJ/mol}$
18. 12.9g of water freeze    {heat absorbed by metal + heat lost by water + heat to freeze water = 0; assume we go to  $0^\circ\text{C}$ : Heat lost by water to  $0^\circ\text{C}$ : Heat =  $(4.184)(50)(0-75) = -15690\text{J}$ ; Heat absorbed by metal to  $0^\circ\text{C}$ : Heat =  $(4)(50)(0 - (-100)) = 20000\text{J}$ ; since the metal will cool the water down to  $0^\circ\text{C}$  and the metal at this time will still be lower than  $0^\circ\text{C}$ , some of the water will freeze  $\rightarrow 20000 - 15690 - 333(\text{x}) = 0$  where  $\text{x} = \text{grams of water that freezes}$ ;  $\text{x} = 12.9\text{g}$ }
19. 15,500J    {Heat = heat to heat ice from  $-20$  to  $0^\circ\text{C}$  + heat to melt ice + heat to heat water from  $0$  to  $100^\circ\text{C}$  + heat to boil water + heat to heat water gas from  $100$  to  $130^\circ\text{C} = (C_{\text{H}_2\text{O}(\text{s})} \times \text{mass}_{\text{H}_2\text{O}(\text{s})} \times \Delta T_{\text{H}_2\text{O}(\text{s})}) + (\Delta H_{\text{fusion}} \times g_{\text{ice}}) + (C_{\text{H}_2\text{O}(\text{l})} \times \text{mass}_{\text{H}_2\text{O}(\text{l})} \times \Delta T_{\text{H}_2\text{O}(\text{l})}) + (\Delta H_{\text{vaporization}} \times g_{\text{ice}}) + (C_{\text{H}_2\text{O}(\text{g})} \times \text{mass}_{\text{H}_2\text{O}(\text{g})} \times \Delta T_{\text{H}_2\text{O}(\text{g})}) = (2.03 \times 5 \times 20) + (333\text{J/g} \times 5) + (4.184 \times 5 \times 100) + (2256\text{J/g} \times 5) + (2.0 \times 5 \times 30) = 15,540\text{J}$ }
20.  $53.8^\circ\text{C}$     {Heat first melts ice; second heats water. Heat to melt ice =  $\Delta H_{\text{fusion}} \times \text{mol}_{\text{ice}}$ ;  $\Delta H_{\text{fusion}} = 6.02\text{kJ/mol}$ ;  $\text{mol ice} = 5\text{g H}_2\text{O} \times (1\text{mol H}_2\text{O}/18\text{g H}_2\text{O}) = 0.278\text{mol H}_2\text{O}(\text{s})$ ; Heat to melt ice =  $6.02\text{kJ/mol} \times 0.278\text{mol} \times (1000\text{J/1kJ}) = 1672\text{J}$ ; Heat left after melting ice = Heat initial - Heat to melt ice =  $5000 - 1672 = 3328\text{J}$ ; now heat is used to heat water:  $q = mC\Delta T$  where  $q = 3328\text{J}$ ;  $m = 5\text{g}$  from ice (now water) +  $10\text{g}$  water =  $15\text{g}$  water;  $C = 4.184\text{J/g}^\circ\text{C}$ ; solve for  $\Delta T$ ;  $\Delta T = q/mC = [3328]/[(15)(4.184)] = 53.84^\circ\text{C} = \Delta T$ ;  $\Delta T = T_f - T_i \rightarrow 53.84 = T_f - 0$  since after the ice has melted the water is at  $0^\circ\text{C} \rightarrow T_f = 53.84 + 0 = 53.84^\circ\text{C}$ }
21. a.  $0.00^\circ\text{C}$     b. 1.58g melt    {Assume all the ice melts and see how much energy this requires: heat ice to  $0^\circ\text{C}$  + melt ice =  $(2.03)(5)(0 - (-10)) + (333)(5) = 101.5 + 1665 = 1766.5\text{J}$  required to do this; cool water to  $0^\circ\text{C}$ :  $(4.184)(15)(0 - 10) = 627.6\text{J}$  released when the water cools; conclusion – not enough heat from water to melt all the ice but enough to heat the ice to  $0^\circ\text{C}$  and melt some of the ice; so set up now becomes:  $q_{\text{heat ice}} + q_{\text{melt ice}} + q_{\text{cool water}}$  with  $T_f = 0$ ;  $(2.03)(5)(0 - (-10)) + (333)(\text{x}) + (4.184)(15)(0 - 10) = 0$ ; solve for  $\text{x}$ ;  $\text{x} = 1.58\text{grams of ice melts}$ }