P C2 Energy - Phase Changes & Temp Changes

Name_______________________Per_____

Important Information
ΔH_{fusion} for solid ↔ liquid = ±334 J/g
$\Delta H_{vaporization}$ for liquid ↔ gas = ±2260. J/g
c for solid water = 2.2 J/gºC
c for liquid water = 4.186 J/gºC
c for gaseous water = 1.7 J/gºC

1. Consider the conversion of 30.0 g of ice at –15.0ºC to steam at 125.0ºC?
   a. What would a sketch of this heating curve (energy diagram) look like?
   b. How much energy is required? Is the energy put in or taken out?
   c. Which part of the sloping portions of the graph would have the steepest slope?
   d. Why is the plateau for vaporization so much longer?

2. Calculate the energy given off when 15.0 g of steam at 100.0ºC condenses and cools down to 30.0ºC.

3. Calculate the energy given off when 15.0 g of liquid water at 100.0ºC cools down to 30.0ºC

4. After calculating #'s 2 & 3, explain why a burn from steam is can be much more severe than a burn from boiling water.

5. Consider holding 15.0 g of ice (assume the ice is at 0ºC) in your hands. Calculate the energy your hand would lose if this ice was melted and then warmed to the temperature of your hand, 30ºC.

6. When 0.850 kJ of energy is added to 10.0 g of ice at 0.0ºC does it all melt?
   a. If yes, to what temperature will the water raise up to?
   b. If no, how much ice is left unmelted?

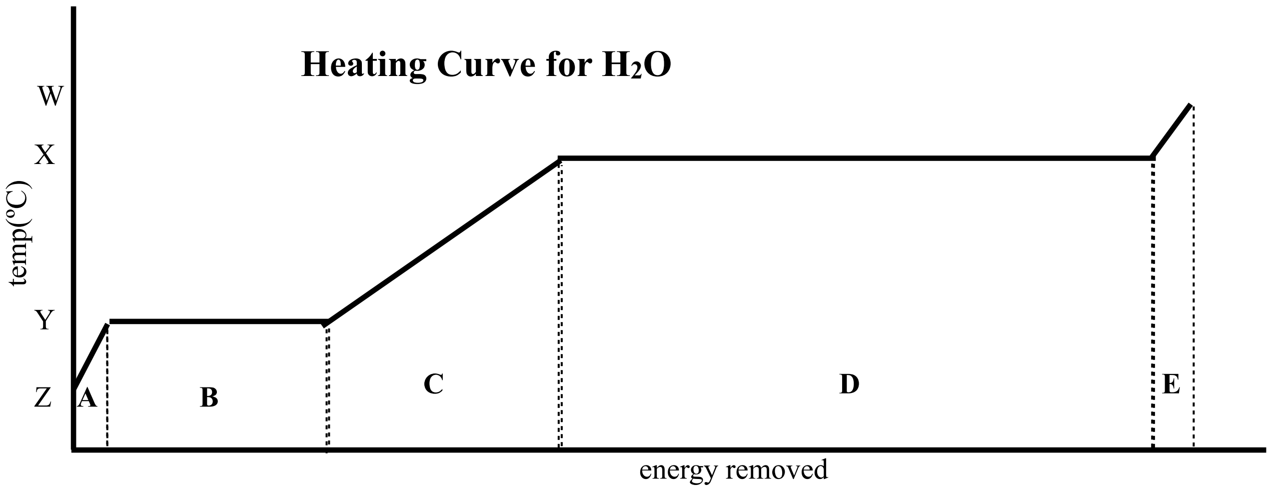
7. What mass of ice, at 0.0ºC would be needed to transform 550.0 g of 37.0ºC water to a cool drink at 0.0ºC?

8. Calculate the mass of ice that must be added in order to cool 5.0 g of water from 20ºC to 0ºC.

9. Calculate what temperature 27.0 g of hot water would need to be at to melt 16.0 g of ice (assume the ice is at 0ºC and the hot water should cool to 0.0ºC)

10. Your 275 g cup of coffee cools down from 75.0ºC to room temperature at 20.0ºC. How much energy in kilojoules was released to the room? (Assume the specific heat capacity of coffee is close enough to liquid water.)

1. Consider the conversion of 30.0 g of ice at –15.0ºC to steam at 125.0ºC?
   a. A sketch of the graph might look something like this.



   b. This is a five part problem using the formula for temperature changes and the formula for phase changes. The energy is put in because the temperature is increased. The total energy **92,800 J** from adding up the five parts of the energy as indicated below. *Significant figures were carried along until rounding off at the end of the problem.*

   A $2.1\frac{J}{gºC}\times 30g\times 15ºC = 945J$

   B $334\frac{J}{g}\times 30g = 10{,}020J$

   C $4.186\frac{J}{gºC}\times 30g\times 100ºC = 12{,}558J$

   D $2{,}260\frac{J}{g}\times 30g = 67{,}800J$

   E $1.7\frac{J}{gºC}\times 30g\times 25ºC = 1{,}275J$

   c. The steepest part of the graph will be the section that represents the phase with the lowest specific heat capacity. Since a lower specific heat capacity will mean that it takes less heat to change the temperature of the substance, its temperature can change more quickly.

   d. The plateau that represents the vaporization (condensation) will be longer for all substances. Since it always requires more energy to break *all* inter-particle forces to vaporize (throw molecules into the air) than it does to break only some inter-particle forces to melt (unglue them from each other).

2. **−38,300 J** total energy given off. Again this is a multi-step problem.

   first calc the energy given off (negative) when the steam condenses $2{,}260\frac{J}{g}\times 15g = -33{,}900J$

   then calc the energy given up to cool down to your hand's temperature $4.18\frac{J}{gºC}\times 15g\times\left(-70ºC\right) = -4{,}395J$

3. **−4,395 J** total energy. This is a single-step problem.

$$4.186\frac{J}{gºC}\times 15g\times\left(-70ºC\right) = -4{,}395J$$

4. Since steam landing on your hand delivers the $\Delta H_{vaporization}$ in order to condense, and *then* delivers the energy to cool from 100ºC to 30ºC the total 38,304 J calculated in #7. Just boiling water on your hand would only deliver the 4,389 J calculated in #8. The extra 33,915 J delivered to your hand packs enough extra energy for a much more substantial burn.

 more Phase Changes & Temp Changes ***ANSWERS***

5. **6,890 J** total energy to your hand. Again this is a multi-step problem.

   first calc the energy absorbed by the melting ice $334\frac{J}{g}\times 15g = 5{,}010J$

   then calc the energy needed to warm it up $4.186\frac{J}{g^{\circ}C}\times 15g\times 30^{\circ}C = 1{,}884J$

   Then add the two heats together to get the total heat absorbed by the ice then water this assumes that the heat absorbed by the water is the same as the heat lost by your hand.

6. **No the ice doesn't all melt, 7.46 g will remain unmelted**.
   first convert 0.850 kJ to 850 J then determine how much ice 850 J will melt

   $850J = 334\frac{J}{g}\times m_{ice}$ m_{ice} = 2.54 g is the mass of ice that can melt,

   then subtraction from 10.0 g will leave you with the mass of ice that did not melt.

7. **255 g of ice would be needed**.
   In this problem you must realize that the liquid water gives up heat as it is cooling from 37ºC to 0ºC. It is giving up its energy to the ice that is melting at 0ºC (assume no temp change for the ice). This allows you to assume that the heat lost by the water is equal to the heat gained by the ice. So set our two equations at the top equal to each other and solve for the mass of ice needed.

   (heat lost by the water) $4.186\frac{J}{g^{\circ}C}\times 550g\times 37^{\circ}C = 334\frac{J}{g}\times m_{ice}$ (heat gained by the ice) Solve for m_{ice} = 255 g

8. **1.25 g of ice would be needed**.
   Just as you did in problem #11, you must realize that the heat lost by the water will equal the heat gained by the ice. This lets us set our two equations equal to each other, and then solve for m_{ice}.

   (heat lost by the water) $4.186\frac{J}{g^{\circ}C}\times 5g\times 20^{\circ}C = 334\frac{J}{g}\times m_{ice}$ (heat gained by the ice) Solve for m_{ice} = 1.25 g

9. **47.4ºC**.
   Just as you did in problem #11, you must realize that the heat lost by the water will equal the heat gained by the ice. This lets us set our two equations equal to each other, and then solve for ΔT.

   (heat lost by the water) $4.186\frac{J}{g^{\circ}C}\times 27g\times \Delta T = 334\frac{J}{g}\times 16g$ (heat gained by the ice) Solve for ΔT = 47.3ºC

   and since the T_f= 0ºC, the the T_{start} must equal 47.3ºC

10. **63.3 kJ**

    $q = mc\Delta T$

    $q = 275g\times 4.186J/g^{\circ}C\times(20-75^{\circ})$