Heating and Cooling Curves

How do you turn ice into steam?

- Duh, like, *heat it?!*
- But, what if I want to know how much heat it would take - say - to turn 100g of solid ice at -10°C to steam at 110°C?

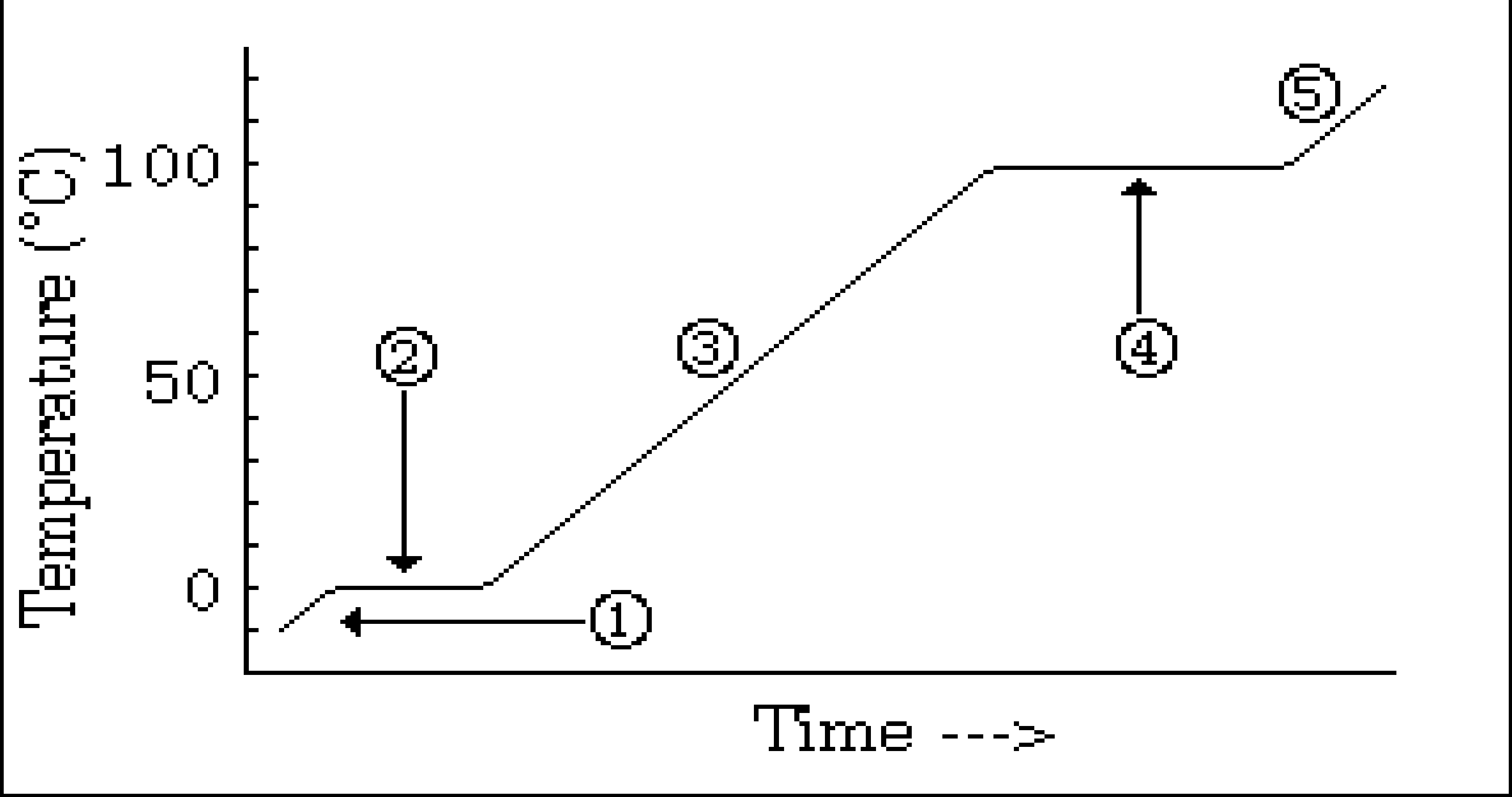
There are actually five steps, each with their own energy calculation

1. raise the T of the ice to 0°C ($q=ms\Delta T$)
2. melt the ice ($q=n\Delta H$)
3. raise the T of the water to 100°C ($q=ms\Delta T$)
4. boil the water ($q=n\Delta H$)
5. raise the T of the steam to 110°C ($q=ms\Delta T$)

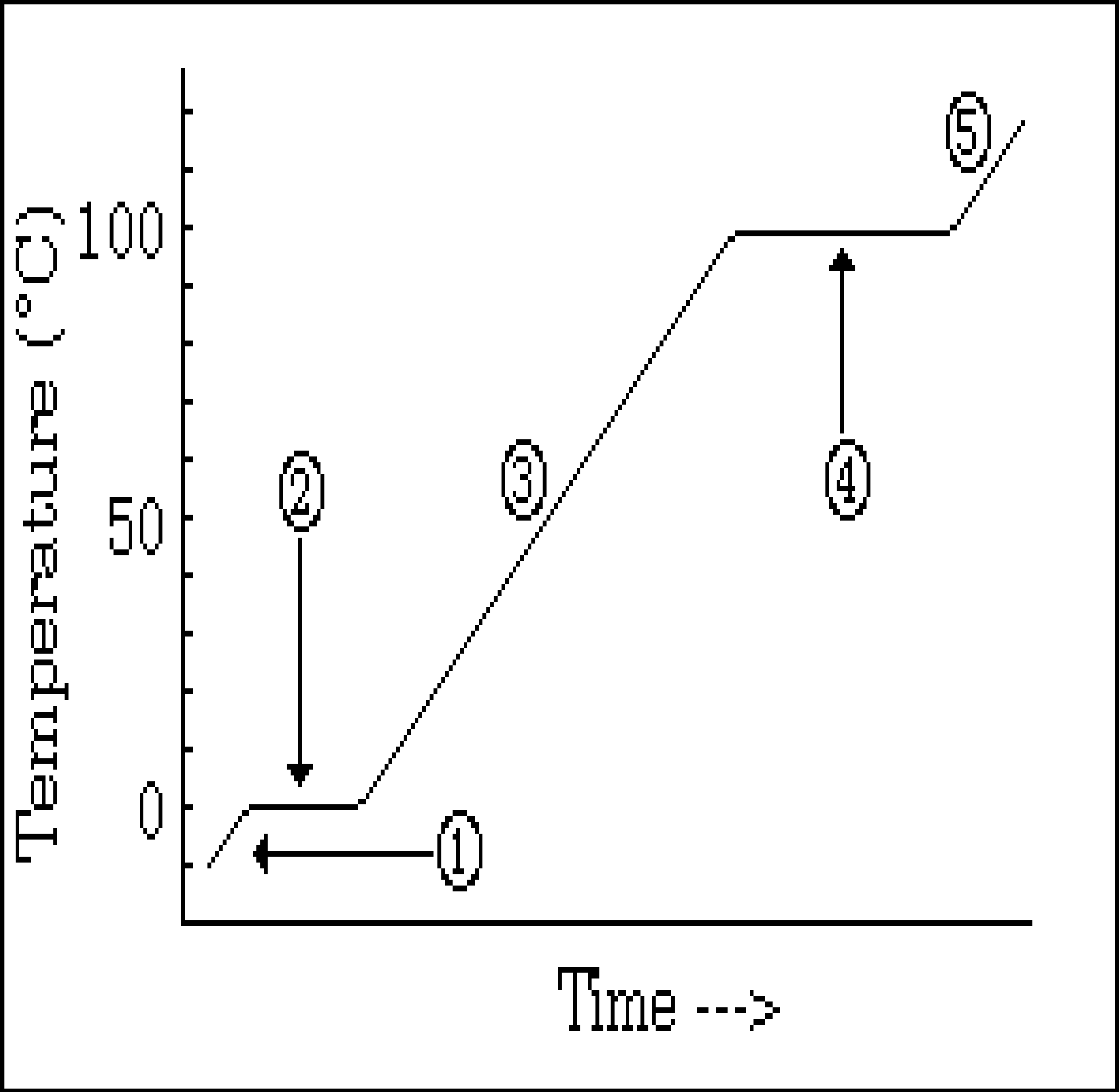
What does this look like graphically?

- We will plot temperature on the y axis, and time on the x axis
- The result is a called a heating curve or a time-temperature graph

The Time-Temperature Graph



- step 1 =
- ↑T ice
- step 2 = melting ice
- step 3 = ↑T water
- step 4 = boiling water
- step 5 = ↑T steam



"Heating" the ice

- Heating a solid without a change in state involves a "ΔT" step using the specific heat of the solid:
- The $s_{ice} = 2.06 J/g^{\circ}C$
- $q = ms\Delta T$
- $q = (100g)(2.06J/g^{\circ}C)(10^{\circ}C)$
- $q = 2060J$
- $q = 2.06kJ$ to "heat" the ice

Melting the ice

- Melting a solid is a change of state, so a "ΔH" not a "ΔT"
  - $q=n\Delta H$
- requires an amount of heat termed the "<u>heat of fusion</u>" (ΔH_{fus})
- the ΔH_{fus} for ice is 6.01kJ/mol

- First, convert 100g of ice into moles (n)
  - 100g x 1 mol/18.02g
  - = 5.55mol
- Next, calculate the amount of heat needed
  - 5.55mol x 6.01kJ/mol
  - = 33.36kJ
- This gives us liquid water at 0°C

Heating the water

- to raise the T of the liquid water to 100°C is a specific heat and "ΔT" problem
  - there is no change of state
- Remember, s_{H2O} = 4.184J/g°C
- q = ms ΔT
- q = (100g)(4.184J/g°C)(100°C)
- q = 41,840J = 41.84kJ
- at this point, what we have is very hot liquid water (100°C)

Boiling the water

- Next up, we must boil the water
- Again, this is a change of state, not a ΔT, so this requires a ΔH amount of heat
- For this process, it is termed the "<u>heat of vaporization</u>" (ΔH_{vap})
- ΔH_{vap} for water is 40.7kJ/mol

- We've already determined that 100g of water is 5.55mol of water
- 5.55mol x 40.7kJ/mol
- = 225.89kJ
- now, what we have is "steam" at 100°C

Heating the steam

- The one step we have left is to "heat the steam up" (change the T) to the 110°C required
- This is a ΔT process, so again, it is a specific heat problem
- $s_{steam} = 2.02 J/g^{o}C$

- $q = ms\,\Delta T$
- $q = (100g)(2.02J/g^{o}C)(10^{o}C)$
- $q = 2020J$
- $q = 2.02kJ$
- now we have steam at $110^{o}C$

How much energy did it take?

- To find the total energy for the process, add the energies required for each step along the way
- Make sure the units on each of the energy terms are the same before adding

- q_{total} =
- 2.06kJ → ↑T of the ice
- + 33.36kJ → melting the ice
- + 41.84kJ → ↑T of water
- + 225.89kJ → boiling water
- + 2.02kJ → ↑T of steam
- = 305.17kJ for the entire process

A few questions…

- Why is the curve "flat" for melting and freezing?
- Which processes involve the most energy?
- How much heat would have to be removed to change 100g of steam at 110°C to ice at -10°C?

Why is the curve "flat" for melting and freezing?

- Changes of state do not involve changes in temperature
- Energy gained is used to separate molecules, or is released if molecules are brought together
- This is a ΔPE process, not a ΔKE
- KE is related to temperature

Which processes involve the most energy?

- Notice that the changes in state require large amounts of energy compared to changing the temperature
- IMF's or bonds are being broken or formed

How much heat would have to be removed to change 100g of steam at 110°C to ice at -10°C?

- These processes are reversible
- 305.17kJ for the entire process would be released

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