Chapter 10 - Solids & liquids

When we heat a solid or a liquid, how do we calculate temperature for that object.

Specific Heat C

The amount of energy required to increase the temperature of 1 g of material by 1°C

Energy units

Cal = 1 cal equals the amount of energy required to raise 1 g of H_2O by 1°C

4.186 joule = 1 Cal

Specific Heat $C = \frac{Cal}{g\,°C}$

	C cal/g°C
$H_2O_{(l)}$	1.01 cal/g°C
$Cu_{(s)}$	0.09197 cal/g°C
$H_2O_{(s)}$	0.4897 cal/g°C
$H_2O_{(g)}$	0.4969 cal/g°C
$Air_{(g)}$	0.2419 cal/g°C

To change the temperature of a cup of tea from 25°C to 100°C would require how much energy if the cup contains 300. mL?

Energy = $mC\Delta T$

- m = mass in g
- C = Specific Heat cal/g°C
- ΔT = Change in temp $(T_f - T_I)$

$m = 300.\,mL \times \frac{1.00\ g}{1\ mL} = 300.\,g$

density = g/mL

$d_{H_2O} = 1.00\ g/mL$

$C = 1.01\ cal/g°C$ (From table or from problem)

$\Delta T = 100°C - 25°C = 75°C$

$E = mC\Delta T$

$= (300.\,g)(1.01\ \frac{cal}{g°C})(75°C)$

(sig figs: 3, 3, 2)

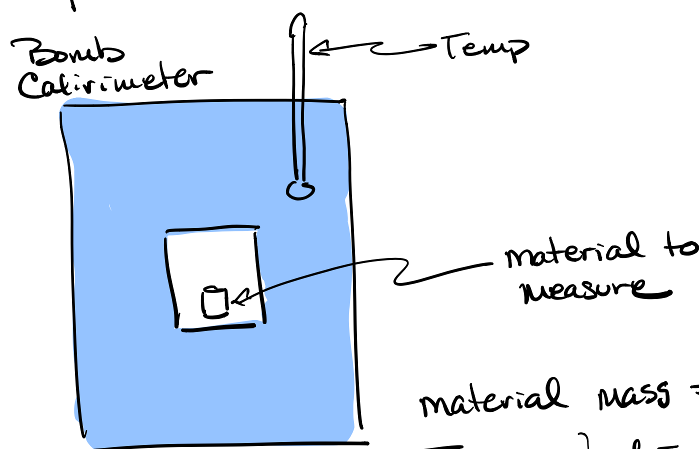
$= 22725\ cal$

$= 23000\ cal$ or $2.3 \times 10^4\ cal$ or $23\ kcal$

calorie = 1 cal = 4.186 Joules

Calorie = food Calorie = 1 kcal = 1000 calories

How are specific Heats measured?



material mass = 100g

Temp initial =

Combust material

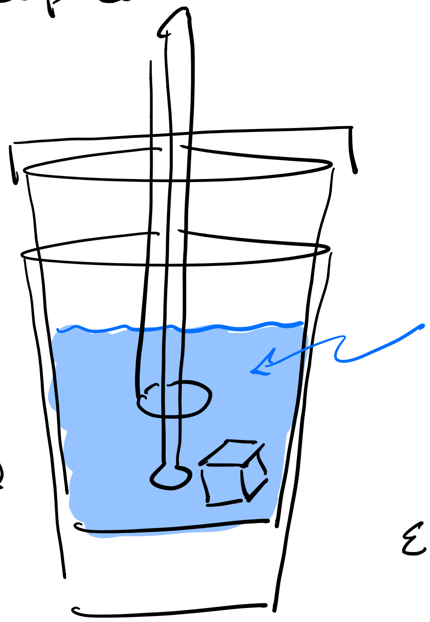
↓

Final Temp =

ΔT = Change in temp
for water

Energy came from Combustion
of material

Coffee Cup Calirimeter

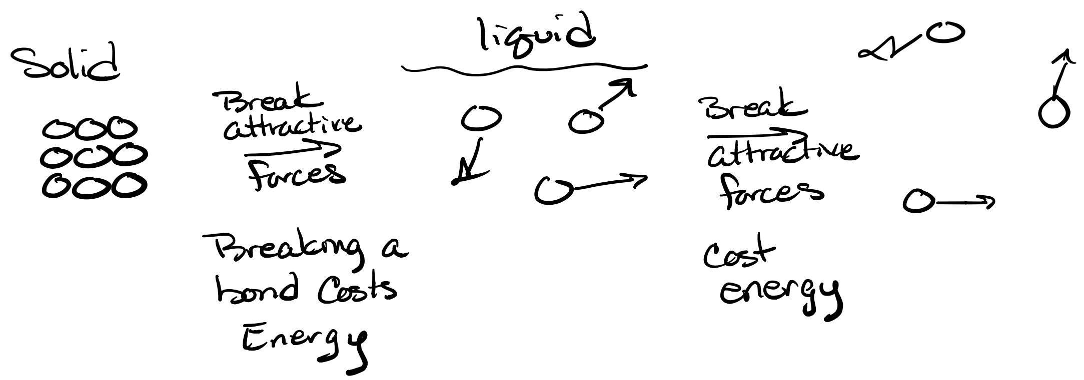


- mass H_2O
- Temp initial
- Add some materal like a block of metal of known mass and temp
- Allow system to come to equilibrium
  ⇒ Temp stops moving
- measure final temp H_2O

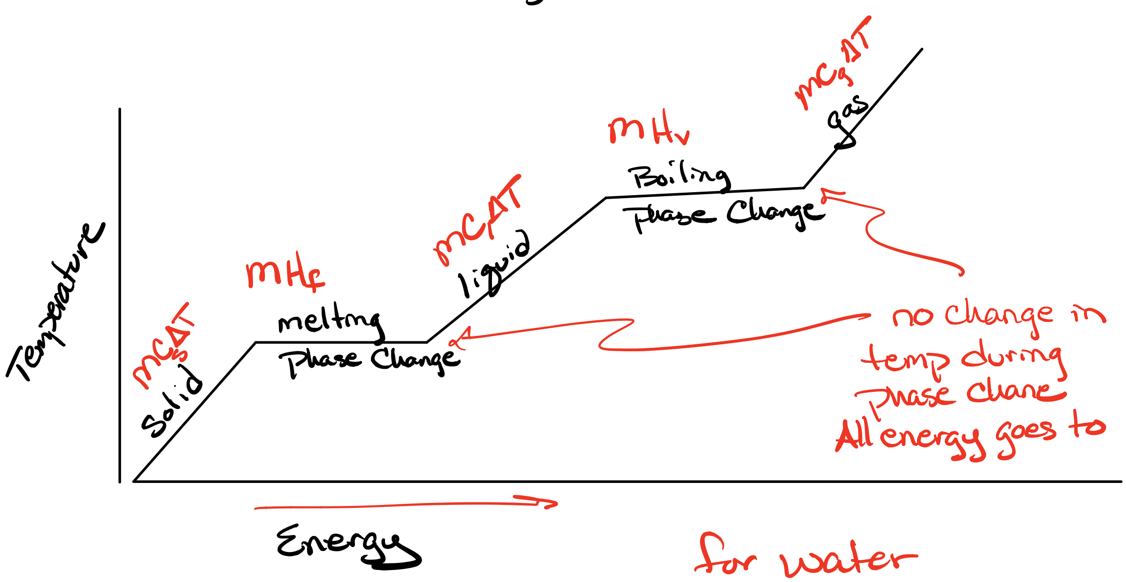
Energy = MCΔT

Phase Changes

Solid ↔ liquid or liquid ↔ gas



Phase Energy Diagram



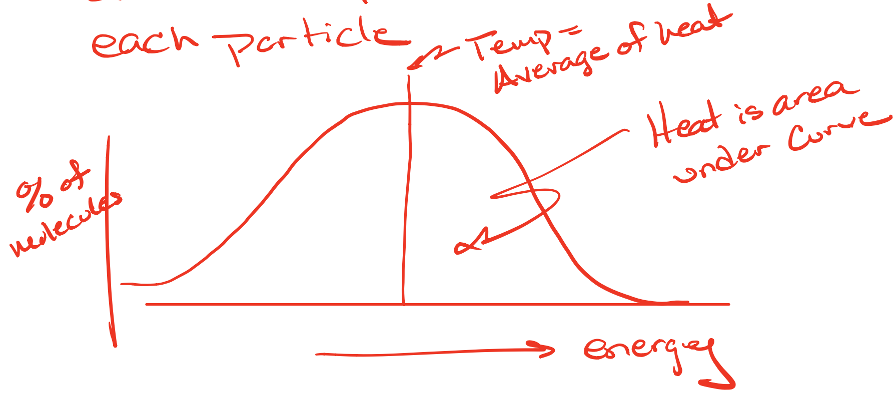
for water

H_f = Heat of fusion 79.7 cal/g

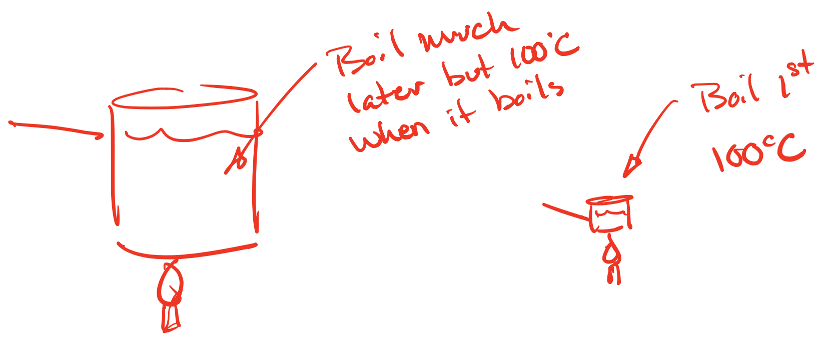
H_v = Heat of Vaporization 540 cal/g

Heat vs. Temp

Heat = amount of energy that is in a system
sum of each particle times the energy of each particle



Temp = Average of all energy



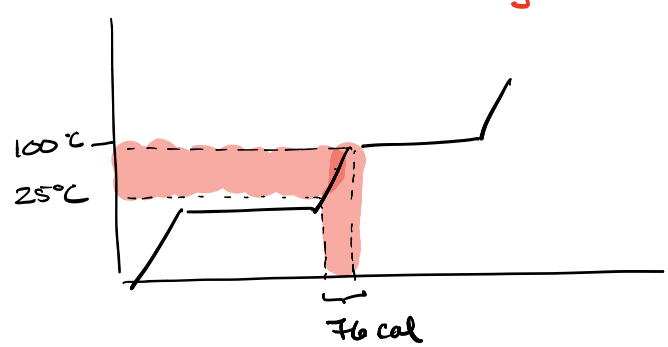
If you spill 100 mL of water on your hand, and that water started at 100.°C. If the water cooled to 25°C, how many calories were released to your hand. $C_{H_2O} = 1.01\ cal/g°C$ $d_{H_2O} = 1.00\ g/mL$

Road Map

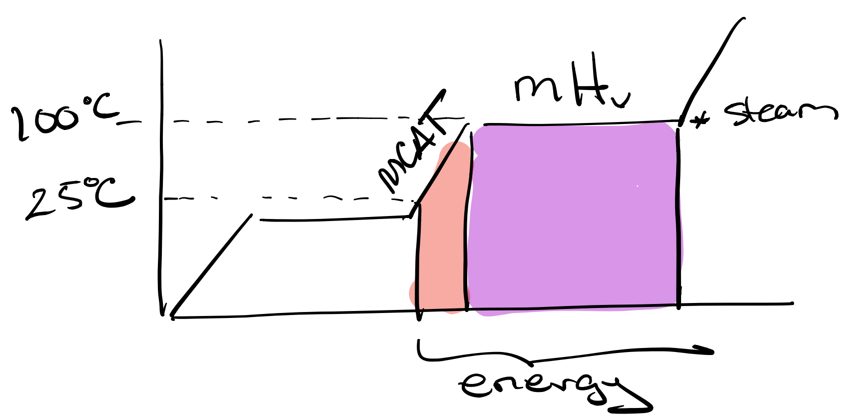
mL H_2O → g H_2O —(mcΔT = E)→ cal released

$$\left(1.00\ mL \times \frac{1.00\ g}{1\ mL}\right)\left(1.01\ cal/g°C\right)\left(100°C - 25°C\right) = 75.75\ cal$$

= **76 cal**



What if it were 1.00 g of steam at 100.°C that contacted your hand & cooled 25°C? How much energy Released?



$$\text{Energy} = mC_{\ell}\Delta T + mH_v$$

$$\left(1.00\,\text{mL} \times \frac{1.00\,\text{g}}{1\,\text{mL}}\right)\left(1.01\,\frac{\text{cal}}{\text{g}^\circ\text{C}}\right)\left(75^\circ\text{C}\right) + \left(1.00\,\text{mL} \times \frac{1.00\,\text{g}}{1\,\text{mL}}\right)\left(540\,\text{cal/g}\right)$$

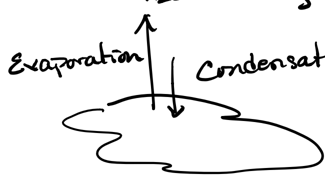
$$= 615.75\ \text{cal}$$

$$= \boxed{616\ \text{cal}}$$

Exaporation

open system

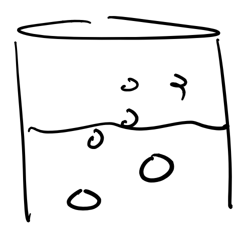
$P_{atm} = P_{H_2O} + P_{gas}$

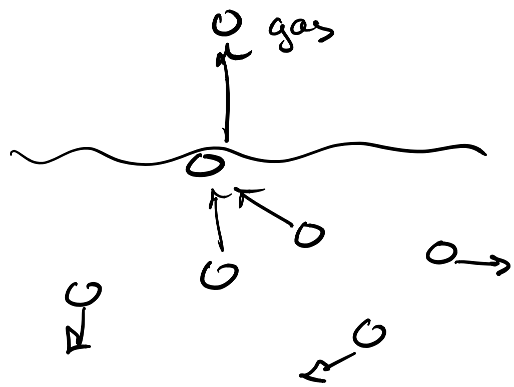


Temp
surface area
pressure

Boiling

$P_{H_2O} = P_{atm}$





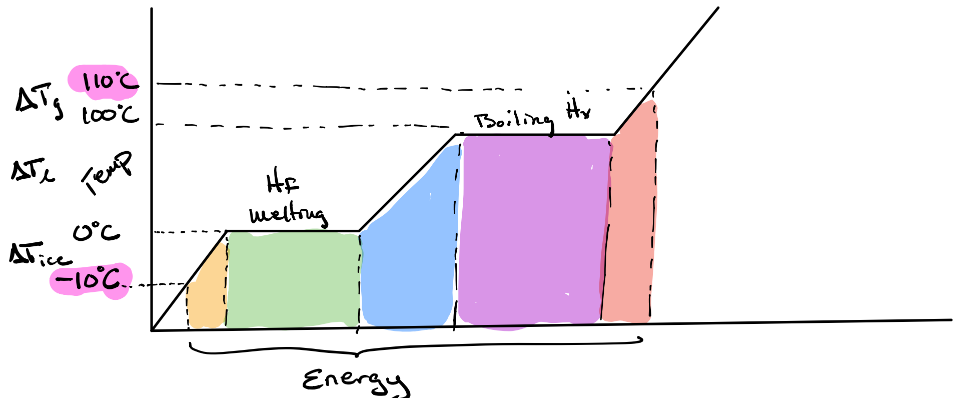
- particle must be on surface
- multiple collisions to impart enough energy

$$\downarrow P_{vap} \propto \frac{1}{\text{attractive forces} \uparrow}$$

How much energy is required to convert a block of ice weighing 200. g from -10°C to steam at 110°C?

$C_\ell = 1.01 \frac{cal}{g°C}$ $C_g = 0.4969 \frac{cal}{g°C}$ $C_s = 0.4897 \frac{cal}{g°C}$

$H_f = 79.7\ cal/g$ $H_v = 540\ cal/g$



$$Energy = mC_s\Delta T_s + mH_f + mC_\ell\Delta T_\ell + mH_v + mC_g\Delta T_g$$

$$= (200.\ g)(0.4897 \tfrac{cal}{g°C})(10°C) + (200.\ g)(79.7\ cal/g) + (200.\ g)(1.01 \tfrac{cal}{g°C})(100°C) +$$

$$(200.\ g)(540\ cal/g) + (200.\ g)(0.4969 \tfrac{cal}{g°C})(10°C)$$

$$= 146113.2\ cal$$

146000 cal = 1.46×10^5 cal

$$(200.\,g)\left(0.4897\,\frac{cal}{g\,^\circ C}\right)(10\,^\circ C) = 979.4\ cal$$

$$(200.\,g)\left(0.4969\,\frac{cal}{g\,^\circ C}\right)(10\,^\circ C) = 993.8\ cal$$

$$(200.\,g)\left(540\,\frac{cal}{g}\right) = 108000\ cal$$

$$(200.\,g)\left(79.7\,\frac{cal}{g}\right) = 15940\ cal$$

$$(200.\,g)\left(1.01\,\frac{cal}{g\,^\circ C}\right)(100\,^\circ C) = 20200\ cal$$

$$+ \quad 146113.2$$

146000 cal

1.46×10^5 cal

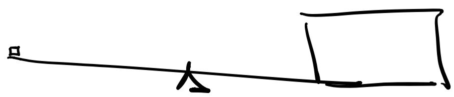
146 kcal

Chapter 13 & 14

↖ equilibrium ↖ acid/Base

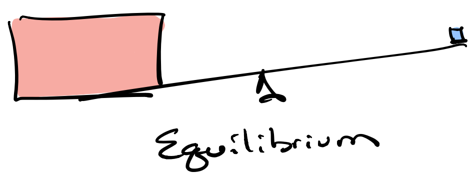
Strong Acid → Dissociates Completely

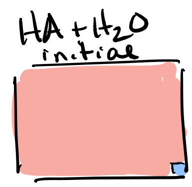
$$HA + H_2O \rightarrow A^- + H_3O^+$$



Weak Acid → Dissociates a little

$$HA + H_2O \rightleftharpoons A^- + H_3O^+$$





Equilibrium → Reaction Quotient

$$aA + bB \rightarrow cC + dD$$

← Chemical Species measured in mols/L

low case = Coefficients

$$Q = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

$[A]$ = mols/L A

(Dynamic) Equilibrium = when the forward Reaction and the reverse reaction have same rate

forward Rate →

← Reverse Rate

$$R_{forward} = R_{reverse}$$

At Equilibrium

$$Q_{eq} = \frac{[C]_{eq}^{c}[D]_{eq}^{d}}{[A]_{eq}^{a}[B]_{eq}^{b}} = K = \text{Equilibrium Constant}$$

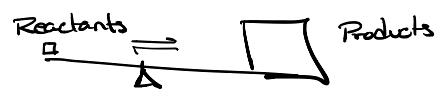
↑ Reaction Quotient measured at any point

$$Q_{\text{at eq}} = K \quad \text{equilibrium Constant}$$

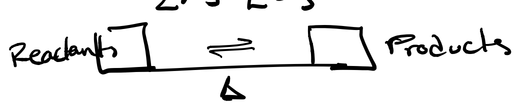
$$aA + bB \rightleftharpoons cC + dD$$

$$K = \frac{[C]^{c}[D]^{d}}{[A]^{a}[B]^{b}}$$

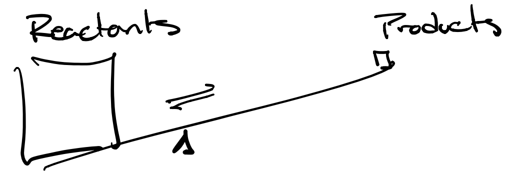
if $K > 1$ $\quad \frac{[C]^{c}[D]^{d}}{[A]^{a}[B]^{b}}$



if $K = 1$ $\quad \frac{[C]^{c}[D]^{d}}{[A]^{a}[B]^{b}}$

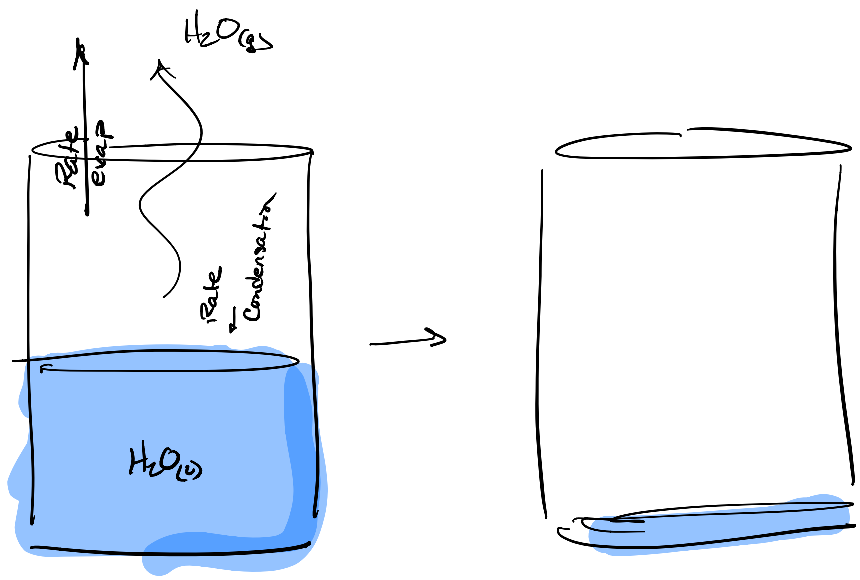


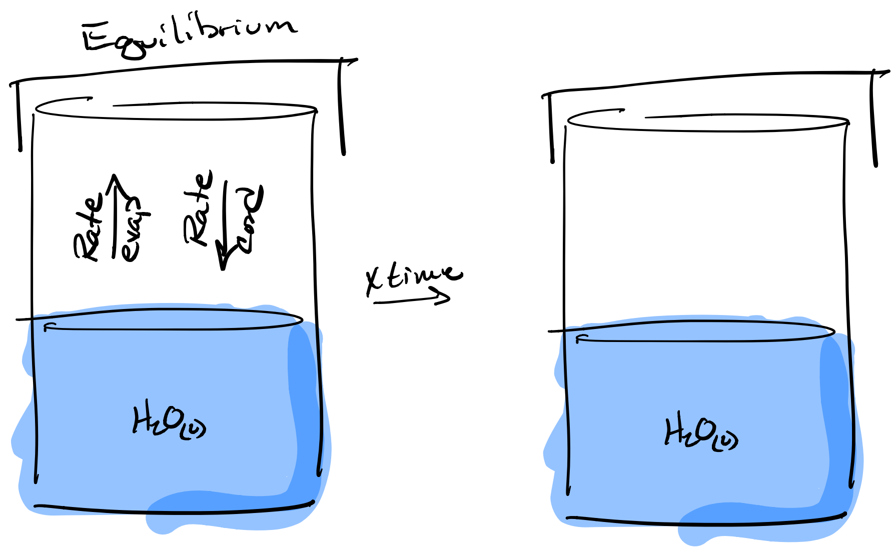
if $K < 1$ $\quad \frac{[C]^c[D]^d}{[A]^a[B]^b}$



$\text{Rate}_{\text{forward}}$ = Rate at which reactants are being converted to products

$\text{Rate}_{\text{Reverse}}$ = Rate at which products are being converted back to reactants.




Equilibrium
Rate evap
Rate cond
x time
$H_2O_{(l)}$
$H_2O_{(l)}$