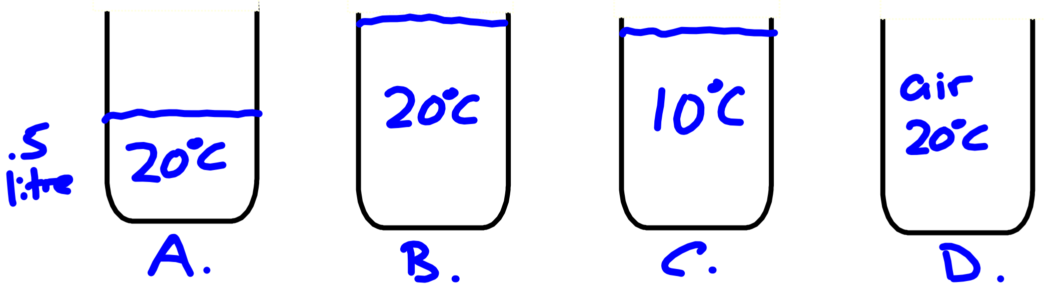
Introduction to Thermal Energy



Jar C has lowest temperature.
Jar B has highest thermal energy content.
Jar D " lowest " " ".
Jar A & C have the same " " ".

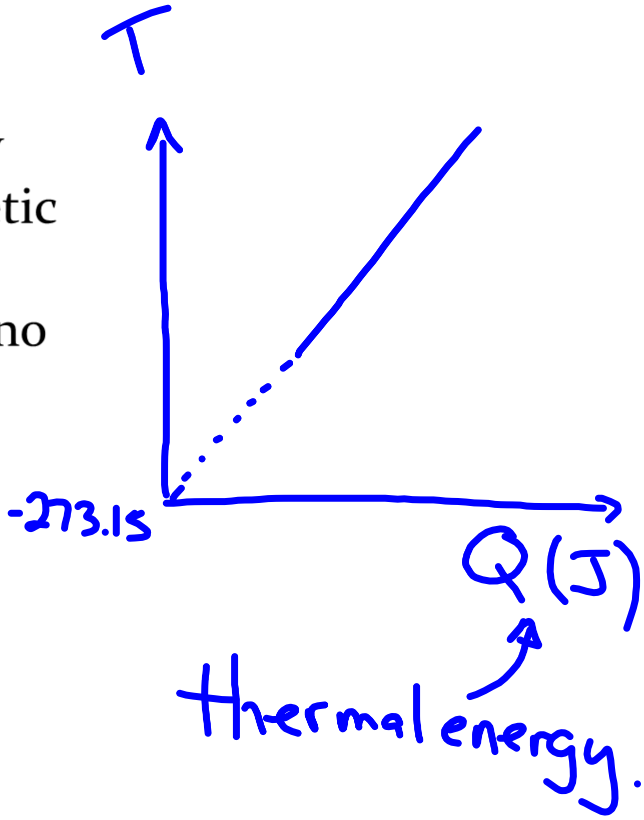
Thermal Energy vs. Temperature vs. Heat

- Thermal energy : is the total kinetic energy of the atoms or molecules of a substance
- Temperature : is a measure of the average kinetic energy of the atoms or molecules of a substance
- Heat : is the transfer of energy from a hot body to a cooler one

Absolute Zero

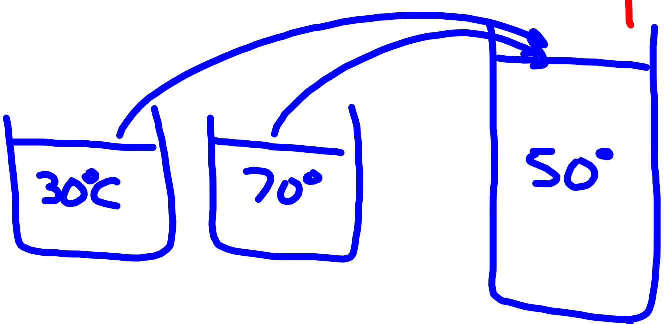
- Absolute zero theoretically occurs when all of the kinetic energy has been removed from a substance (there is no energy remaining).

0° Kelvin = -273.15° C



2nd Law of Thermodynamics

Thermal Equilibrium - thermal energy is always transferred from an object at a higher temperature to an object at a lower temperature. (an isolated system will always progress towards thermal equilibrium). same temperature.



mixed together.

Specific Heat Capacity

Definition :

specific heat capacity of a material is the amount of energy that must be added to raise the temperature of 1 kg of material by 1 °C.

Formula:

mass (kg)

change in temperature (°C)

$$Q = mc\Delta T$$

energy (J)

specific heat capacity (J/kg°C)

Specific Heat Capacity of Common Substances

Substance	Specific Heat Capacity (J/(kg°C)		Substance	Specific Heat Capacity (J/(kg°C)
aluminum	920		water	4180
glass	840		ice	2100
iron	450		steam	2100
silver	240		alcohol	2500
lead	130		human body	3470
air	~1000			

Practice problems assume 2 sig digs in answer!!

1. How much energy is required to heat up 300 g of water from 10°C to 23°C?

2. A 55 kg person generates about 6.5×10^5 J of energy (one hour brisk walk) – how much would our body temperature raise if we did not have an efficient temperature regulating systems?

3. A 200g piece of iron at 350°C is inserted into 300g of water at 10°C to cool it quickly. Determine the final temperature of the water and the iron.

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1. G: $m = 0.300\,kg$

$c = 4180\,J/kg°C$

$\Delta T = 23 - 10 = 13°C$

R: Q

S: $Q = mc\Delta T$

$= (0.300\,\cancel{kg})(4180\,J/\cancel{kg°C})(13\,\cancel{°C})$

$= 16000\,J$

2. G: $m = 55\,kg$

$c = 3470\,J/kg°C$

$Q = 6.5 \times 10^5\,J$

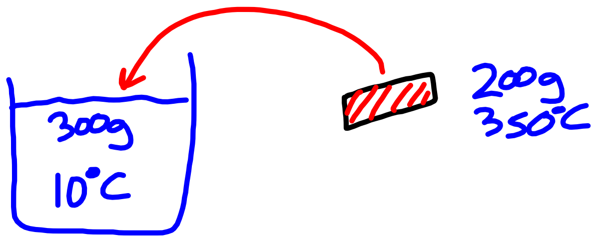
R: ΔT

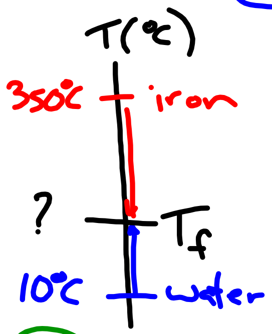
S: $Q = mc\Delta T$

$$\Delta T = \frac{Q}{mc} = \frac{6.5 \times 10^5\,J}{55\,kg \cdot 3470\,J/kg°C}$$

$= 3.4°C$

3.





$\Delta T_{iron} = T_f - 350$

Total energy remains same.

energy lost by iron + energy gained by water = ϕ

$Q_{iron} + Q_{water} = \phi$

$mc\Delta T_{iron} + mc\Delta T_{water} = 0$

$(0.2)(450)\Delta T_{iron} + (0.3)(4180)\Delta T_{water} = 0$

$90\,\Delta T_{iron} + 1254\,\Delta T_{water} = 0$

T_f - final temp

$\Delta T_{iron} = T_f - 350$

$\Delta T_{water} = T_f - 10$

$90(T_f - 350) + 1254(T_f - 10) = 0$

$90T_f - 31500 + 1254T_f - 12540 = 0$

$1344T_f - 44040 = 0$

$T_f = \dfrac{44040}{1344}$

$= 33°C$

$\Delta T_w = 23°$
$(33 - 10)$

$\Delta T_{iron} = -317°$
$(33 - 350)$

Last Days Homework

Handout #6

6. A cold piece of silver (500g) at -75°C and a hot piece of iron (500g) at 85°C are both inserted into a 2.0 litre glass of water at 30°C. What would the final temperature of these three objects be after they have reached thermal equilibrium?

Total Change in Energy = 0.

$$Q_{silver} + Q_{iron} + Q_{water} = 0$$

$$mc\Delta T_{silver} + mc\Delta T_{iron} + mc\Delta T_{water} = 0$$

$$(0.5)(240)\Delta T_{silver} + (0.5)(450)\Delta T_{iron} + (2)(4180)\Delta T_{water} = 0$$

$$120\,\Delta T_{silver} + 225\,\Delta T_{iron} + 8360\,\Delta T_{water} = 0$$

$$120(T_f - (-75)) + 225(T_f - 85) + 8360(T_f - 30) = 0$$

$$120T_f + 9000 + 225T_f - 19125 + 8360T_f - 250800 = 0$$

change Δ = final − initial = $T_f - T_i$ (final, initial)

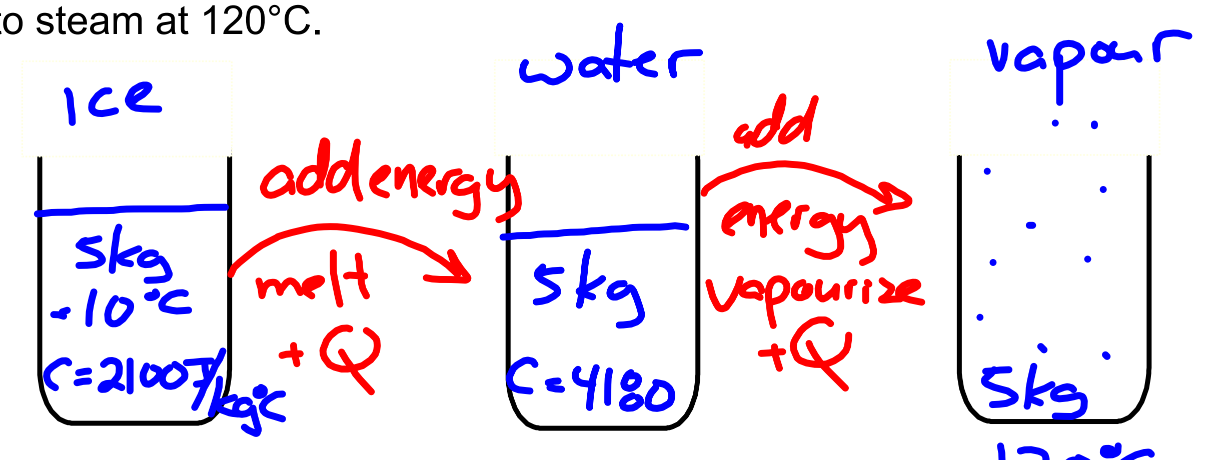
$$8705T_f - 260925 = 0$$

$$T_f = \frac{260925}{8705} = 29.974$$

$$\therefore T_f = 30°C.$$

Thermal Energy - Phase Changes

Calculate the energy required to heat 5.0 kg ice initially at -10°C to steam at 120°C.



$Q = mc\Delta T$

$\Delta T = +10$

$Q = mc\Delta T$

$\Delta T = +100$

120°C

$c = 2100$

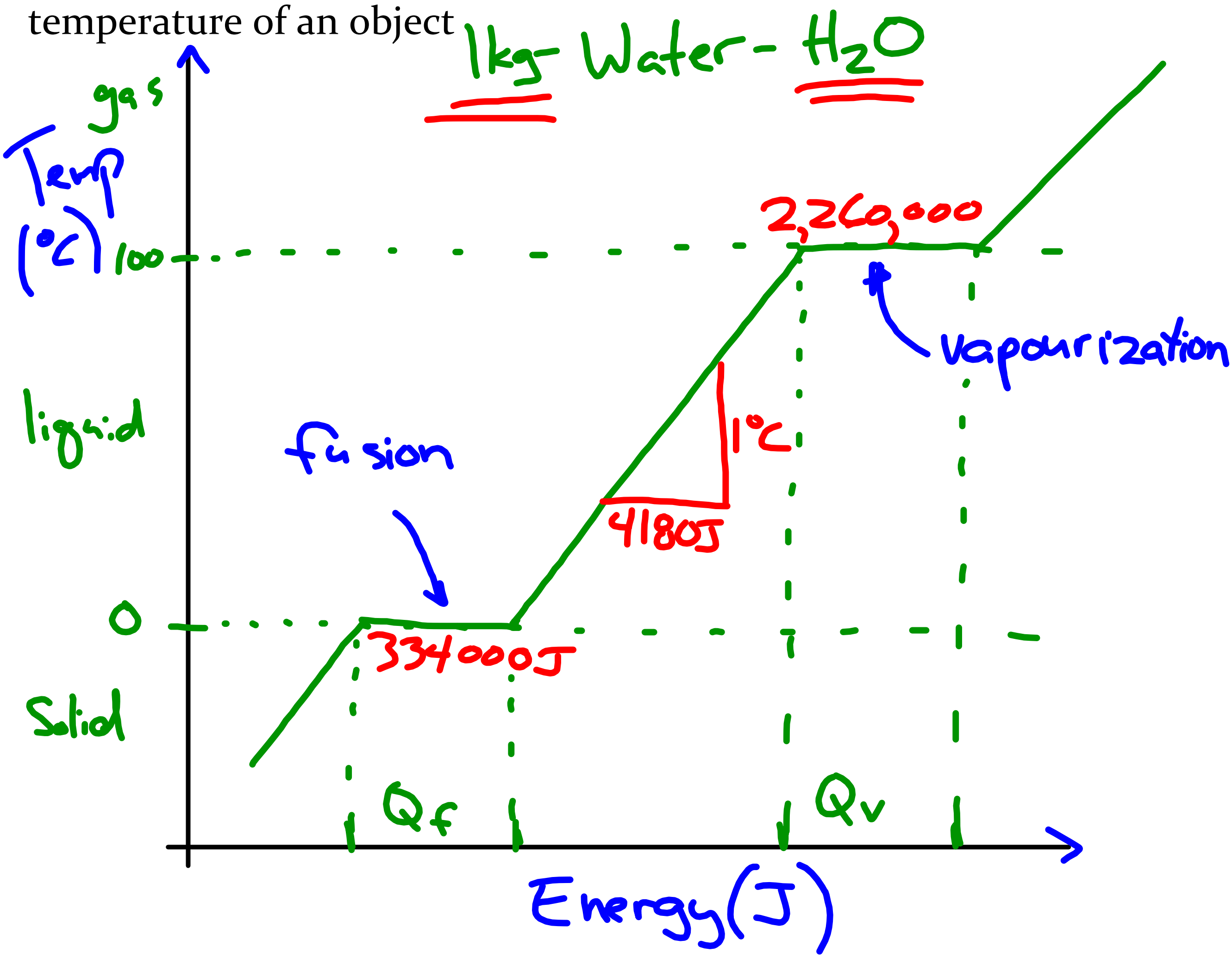
$Q = mc\Delta T$

$\Delta T = +20$

	Substance	Specific Heat Capacity (J/(kg°C)
	water	4180
	ice	2100
	steam	2100

Heating Diagrams

used to show the heat (energy) required to change the temperature of an object



Latent Heat of Fusion:

L_f

Heat energy required to convert 1 kg of material from solid to liquid.

$$Q_f = mL_f$$

$$L_{f\,\text{water}} = 334000\ \text{J/kg}.$$

Latent Heat of Vaporization:

L_v

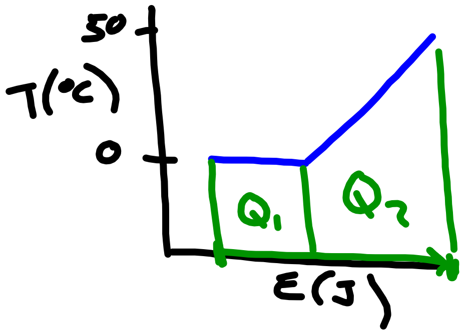
Heat energy required to convert 1 kg of material from liquid to gas.

$$Q_v = mL_v$$

$$L_v = 2,260,000\ \text{J/kg}$$

Practice Problems:

1. Draw a heating diagram and calculate how much energy is required to change 2.0 kg of ice at 0°C to water at 50°C?



$Q_{TOT} = Q_1 + Q_2$

$= mL_f + mc\Delta T$

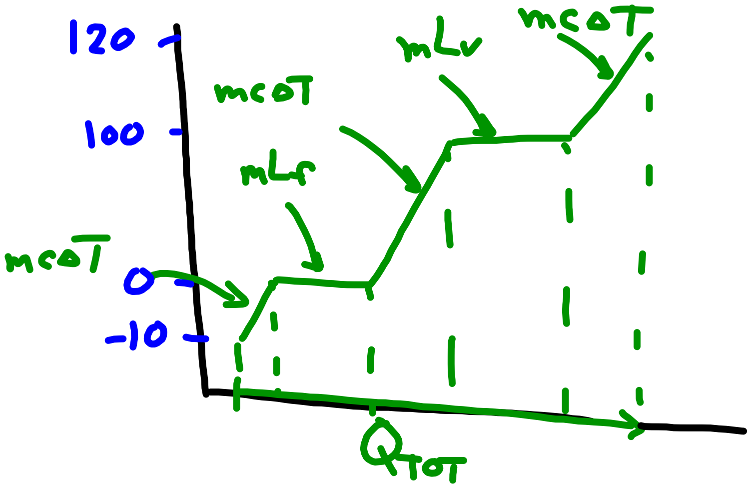
$= (2.0)(334000) + (2.0)(4180)(50)$

$= 1086000$ (1,100,000)

$= 1.1 \times 10^6$ J

Practice Problems:

2. Draw a heating diagram and calculate how much energy is required to change 5.0kg of ice at -10°C to steam at 120°C.

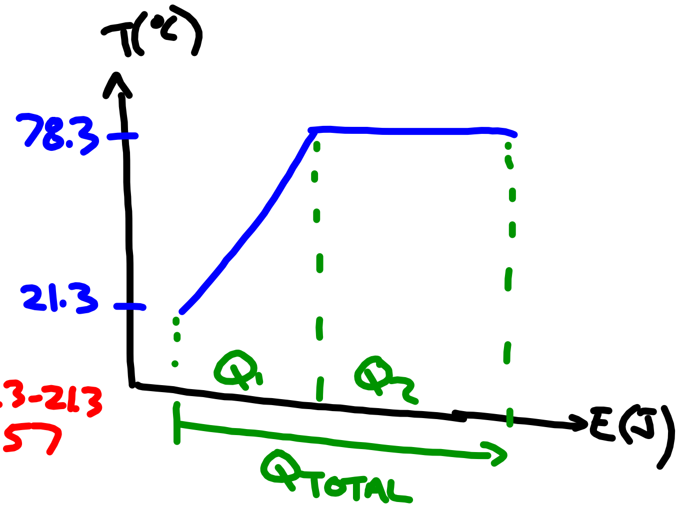


$$Q_{TOT} = \underbrace{mc\Delta T}_{\text{ice}} + mL_f + \underbrace{mc\Delta T}_{\text{water}} + mL_v + \underbrace{mc\Delta T}_{\text{steam}}$$

5. Draw a heating diagram and calculate the energy required to convert 236 grams of ethyl alcohol from room temperature (21.3°C) to vapor.

Freezing Point (°C)	~~-114~~
Boiling Point (°C)	$\boxed{78.3}$
Latent Heat of Fusion (J/kg)	1.1×10^5
Latent Heat of Vaporization (J/kg)	8.6×10^5
Specific Heat Capacity (as a liquid) (J/kg·°C)	2,460

$\Delta T = 78.3 - 21.3$
$= 57$



$Q_{TOTAL} = Q_1 + Q_2$

$= mc\Delta T + mL_v$

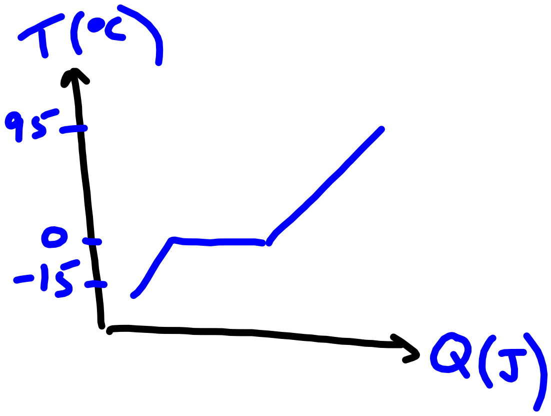
$= (0.236)(2460)(57) + (0.236)(8.6\times10^5)$

$= 240\,000\ J$ $\quad 240\ kJ$

$2.4\times10^5\ J$

Practice Problem - The Cost of Energy

How much would it cost to heat up 0.5 kgs of ice at -15°C to 95°C if energy cost 9.5 cents per kWh?



ice $c = 2100\ J/kg°C$
water $c = 4180\ J/kg°C$
$L_f = 334000\ J/kg$
$L_v = 2260000\ J/kg$

$$Q_{tot} = Q_{\substack{heat\\ice}} + Q_{\substack{melt\\ice}} + Q_{\substack{heat\\water}}$$

$$= (.5)(2100)(15) + (0.5)(334000) + (.5)(4180)(95)$$

$$= 381300\ J$$

$9.5\ ¢/kW \cdot h$

$kW \cdot h \rightarrow$ energy

1 hour = 60x60

$$1 kW \cdot h = 1000 W \cdot h$$

$$= 1000\ J/\cancel{s} \cdot 3600 \cancel{s}$$

$$\boxed{1 kW \cdot h = 3,600,000\ J}$$

Cost = usage × rate

$$= 381300 \times \frac{9.5¢}{3,600,000\ J}$$

$$= 1.0¢$$

Homework / Practice

Text Book
page 152 #10, 11, 13, 14, 15, 16
page 153 #2, 4, 6, 8

handouts :
1. Practice Problems - Heat, Temperature & Thermal Energy
2. Latent Heat of Fusion and Vapourization