Lecture 6
April 14, 2011

Mixing Problems and solution An extra comment re copper

A 2 kg block of copper at 90C° is dumped into 2 gallon bucket of water at 20 C°. What is the final temperature of the water?

Answer: First convert 2 gallons = 2x 3.78 Litre = 7560 cm³. Its weight is 7560 gm since density is 1 gm/cm³
Next use the formula given in last slide.

$$T_f = \frac{M_1 c_1 T_1 + M_2 c_2 T_2}{M_1 c_1 + M_2 c_2}$$

In applying this:
M1= 2kg ; c1= .27 kJ/kg C°; T1= 90 C° Copper
M2= 7.56kg; c2= 4.2 kJ/kg C°; T2= 20 C° Water

Note added:
In class I used a wrong value of c for copper and hence got a large change of temperature. In fact 2 kg copper makes very little difference to the final temperature.

Hence T_f = 21.7C°

Copper density: 8.9 gm/cm³

Hence specific heat can be expressed in two ways
Copper: .27 J/gm or .27J/gm x 8.9 gm/cm³ = 2.4 J/cm³
Compare water 4.2J/cm³ = 4.2J/gm.
Indeed our intuition is right: Copper has comparable specific heat if we reckon it per cm³

We cannot talk of Q (but only of ΔQ) then
What is T anyway? (microscopic level)

T is definable from Maxwell Boltzmann theory of gases/liquids

Thermal agitation. Here Boltzmann's constant makes its appearance.

$$T = \frac{m}{3k_B}\langle v^2 \rangle$$

$$k_B = 1.38 \times 10^{-23} J/K$$

http://www.chm.davidson.edu/vce/kineticmoleculartheory/basicconcepts.html

Changes of phase and Latent heat

Solid state (ice, rock)
Liquid state (water, lava)
Gas state (vapour, dispersed ash)

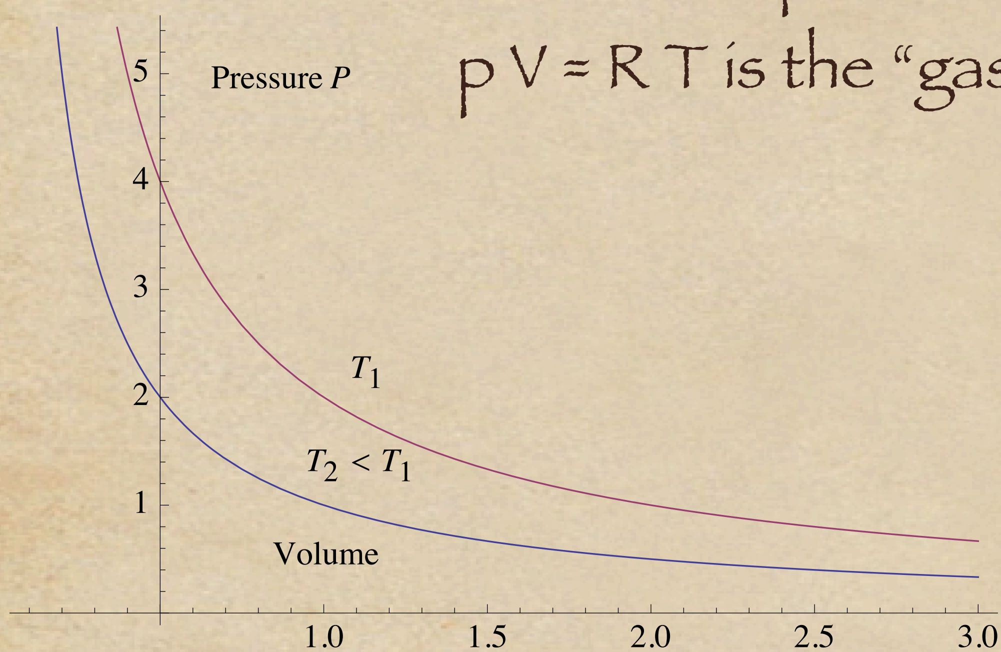
The same matter, in different conditions exhibits different phases. Converting from one phase to another either absorbs energy from environment (endothermic) or adds energy to the environment (exothermic).

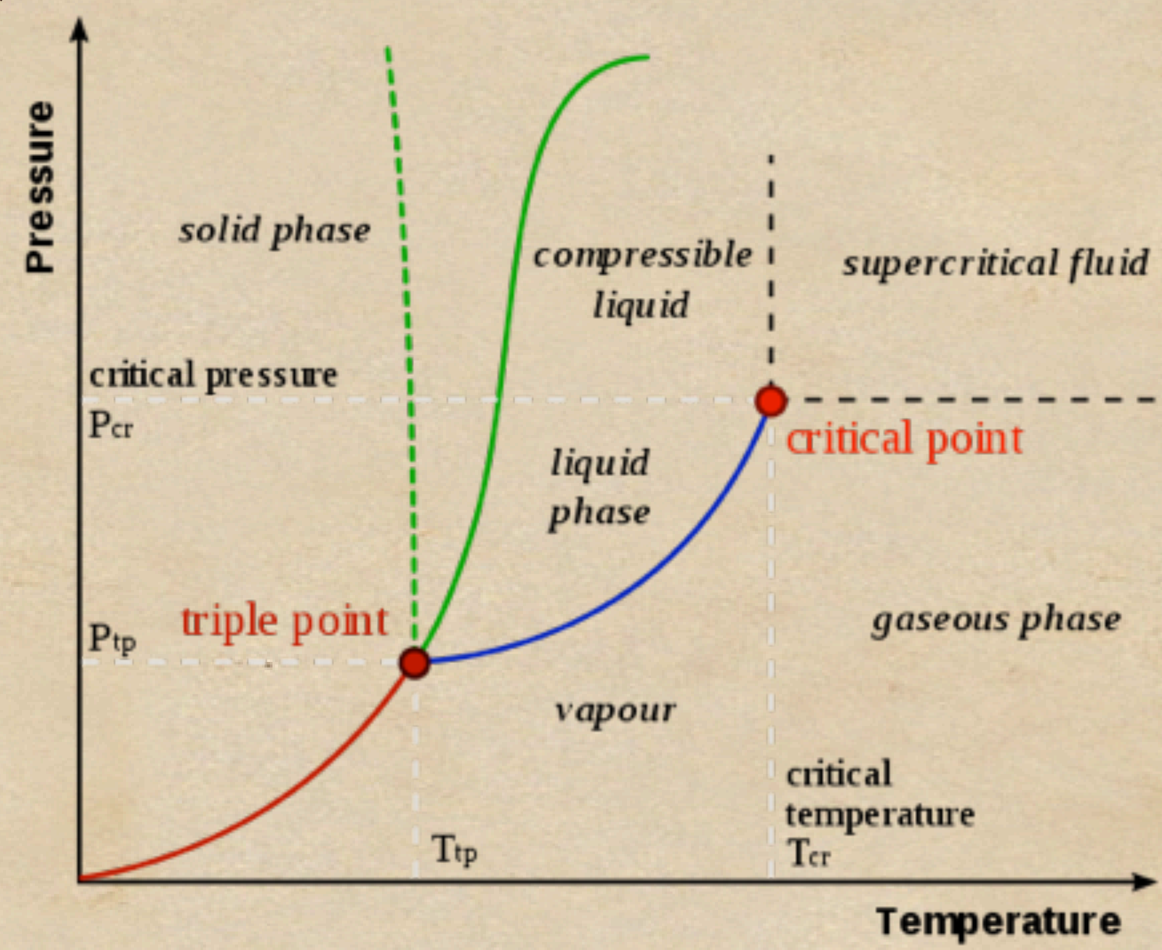
ice melts	$\Delta Q<0$	Need to supply heat from neighbourhood
water freezes	$\Delta Q>0$	warms the neighbourhood
water evaporates	$\Delta Q<0$	Need to supply heat from neighbourhood
vapour condenses	$\Delta Q>0$	warms the neighbourhood

Phase diagram and latent heat

Gases compress. T also compresses

p V = R T is the "gas law"



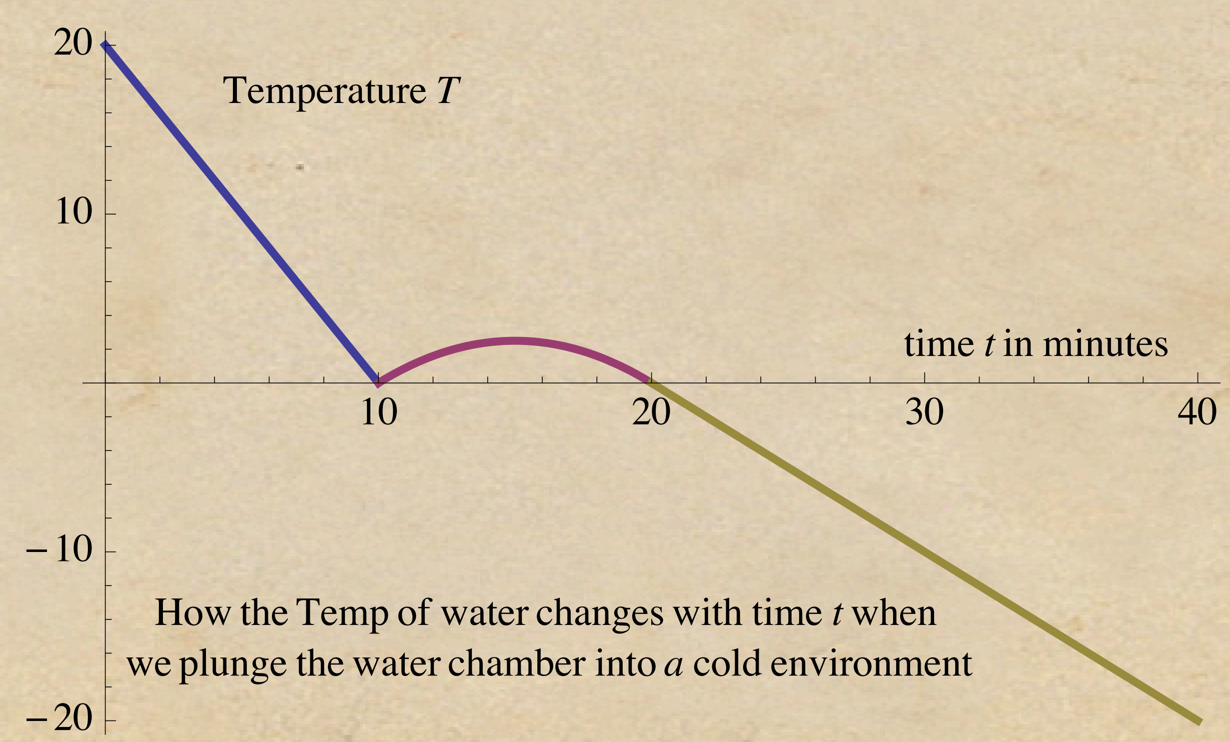


When thermal energy is withdrawn from a liquid or solid, the temperature falls. When thermal energy <u>is</u> added to a liquid or solid, the temperature rises. However, at the transition point between solid and liquid (the melting point), extra energy is required (the heat of fusion).

In going from liquid to solid (freezing), the molecules of a substance become arranged in a more ordered state. Since the solid is more ordered, it has a lower energy and the excess energy is released- exothermic in this process.

In going from solid to liquid (melting), the molecules of a substance become arranged in a less ordered state. Heat energy is needed to "break" the solid's crystalline order hence absorbs i.e. endothermic. In both cases change of state occurs at fixed temperature.

The heat of fusion can be observed by measuring the temperature of water as it freezes. If a closed container of room temperature water at 20 ^{0}C is plunged into a very cold environment (say –20 °C), the temperature will fall steadily until it drops just below the freezing point (0 °C). The temperature then will rebound and hold steady while the water crystallizes. Once the water is completely frozen, its temperature will fall steadily again.

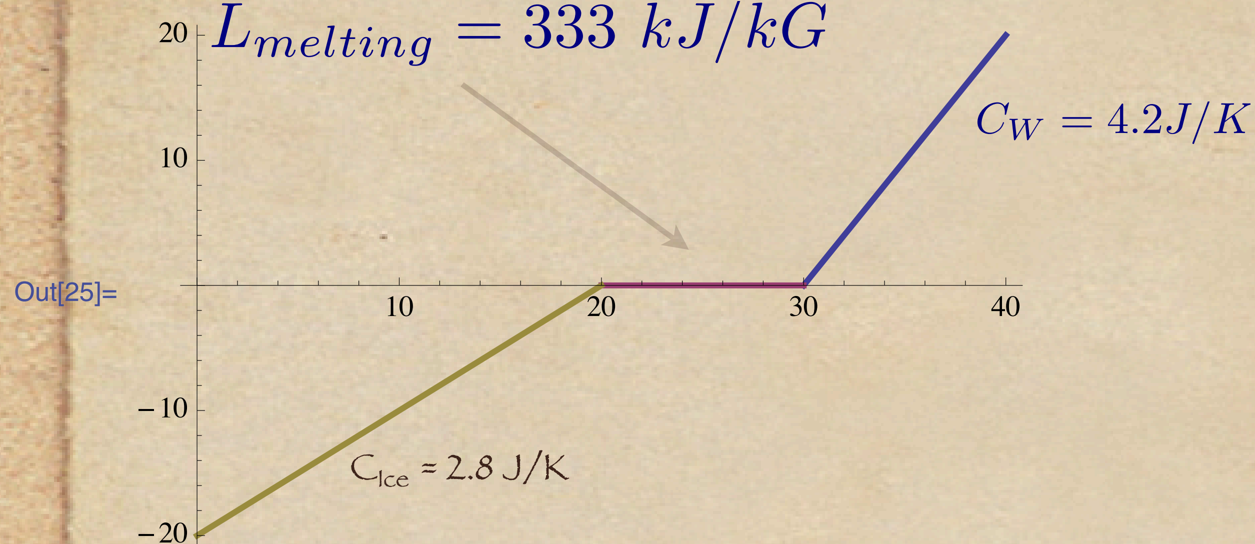


Latent heat is associated with melting and also with vaporizing (boiling)
also sublimation

$$\Delta Q = M\ L$$

Melting of ice: endothermic (needs heat)

$$L_{melting} = 333\ kJ/kG$$



Boiling of water into Steam!
also endothermic

$$L_{vapourizing} = 2.25\ MJ/kG$$

$$\Delta Q = c\ M \Delta T \quad (Specific Heat)$$
$$\Delta Q = M\ L \quad (Latent Heat)$$

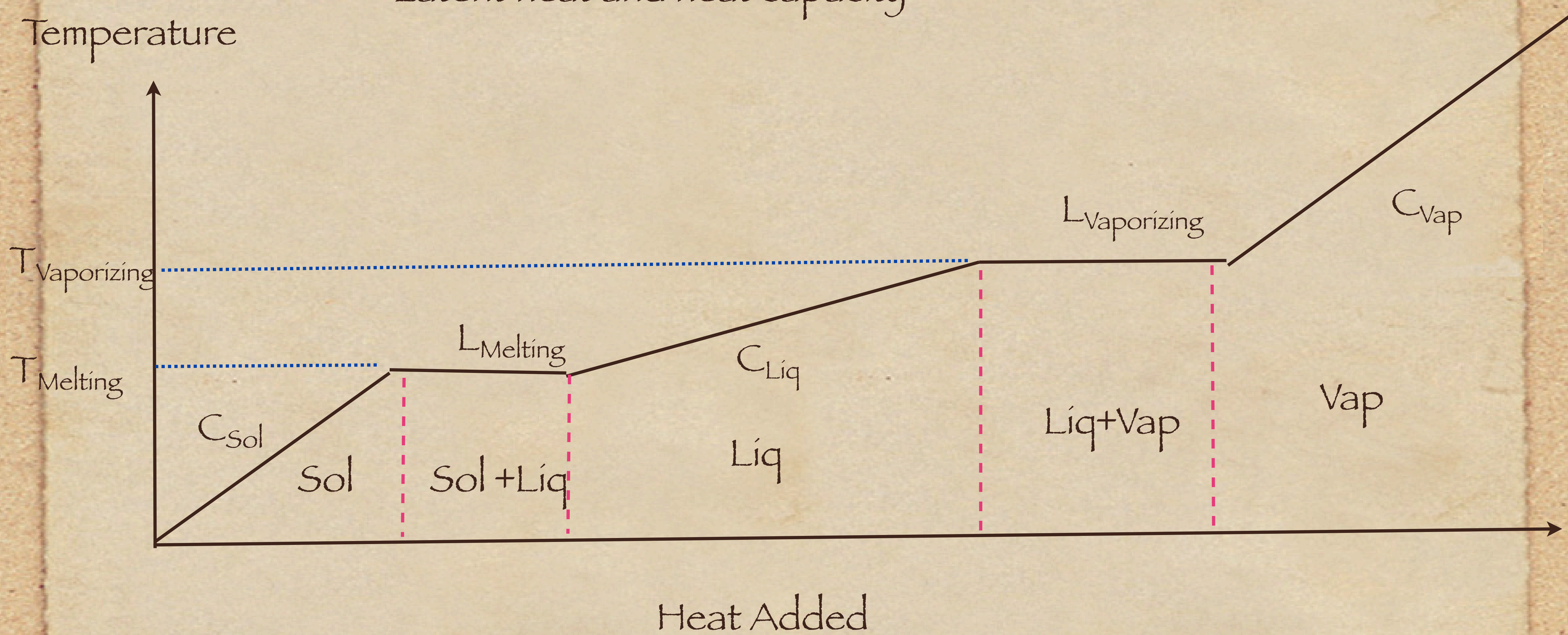
Heat change = Heat due to change of temperature
+ Heat due to phase change

Mixing problems:

Energy is conserved

Phase changes (transitions)

Latent heat and heat capacity



Examples

$$\Delta Q = M\ L$$

$$L_{melting} = 333\ kJ/kG$$

a) 5 kG water freezes: how much heat does it generate in the process? and what is the temperature after freezing?

Using given formula: ΔQ = 5 kG x 333 kJ/kG = 1665 kJ = 1.665 MJ

b) 10 kG ice melts: How much heat does it absorb from the environment?

Answer= 3.333 MJ

c) 10 kG water boils: how much heat does it require?

$$L_{vapourizing} = 2.25\ MJ/kG$$

Answer= 22.5 MJ (Note the much bigger scale)

Summarizing the difference between
Specific heat versus Latent Heat

$\Delta Q = M\,c\,\Delta T$

(State is fixed but T changes)

$\Delta Q = M\,L$

(T is fixed but state changes)

Example combing the two

Find the heat needed to heat 10 kG water at 90°C to steam at 110°C

A) There is heating of water from 90 to 100 C,
B) change of state to steam at 100 C
C) heating of steam from 100 C to 110 C

Data given: Latent heat for boiling 2.25 MJ/kG
Specific heat of water 4.2 kJ/kG
Specific heat of steam 1.996 kJ/kG
Specific heat of ice 2.18 kJ/kG

$\Delta Q = Q_a + Q_b + Q_c$

$Q_a = 420$ kJ, $Q_b = 22.5$ MJ, $Q_c = 199.6$ kJ

Problem in HW3

A) An unknown amount of water at 20 C is mixed with 6 ice cubes at 0C, each with weight 30 grams. The mixture becomes cold water at 5 C. what is the weight of the total mixture?

B) A shot of lead of unknown mass is dropped into 1 litre of water at 30 C and is just hot enough to convert the water to steam at 120 C in equilibrium with the shot. Calculate the mass of the shot. Specific heat of lead = .13 kJ/kG

Process:

1) Locate the appropriate formulas- make sure you have all the needed ones.
2) Identify the object required for answering the question and give it a symbolic name - e.g. "x" kGs in the above problems.
3) Using "x" in the formulas, set up an equation where the unknown is on the LHS and the rest on RHS. Here you have to use (1)
4) Solve for x!! :-)