Warm Up

1675 kJ of energy is added to 8.90 L of water at 18.0°C. What is the new temperature of the water?

$q = 1675\,kJ$

$V = 8.90\,L$

$T_i = 18.0°C$

$T_f = ?$

$C = 4.19 \frac{kJ}{L \cdot °C}$

$q = mC\Delta T$ $\quad \frac{q}{mC} = T_f - T_i$

$q = VC(T_f - T_i)$

$1675\,kJ = (8.90\,L)\left(4.19 \frac{kJ}{L \cdot °C}\right)(T_f - 18.0°C)$

$\frac{1675\,kJ}{(8.90\,L)\left(4.19 \frac{kJ}{L \cdot °C}\right)} = T_f - 18.0°C$

$44.9°C = T_f - 18.0°C$

$T_f = 62.9°C$

5. $6.5 \times 10^3 KJ$

$$\frac{2 \text{ mol } H_2O}{571.6 KJ}$$

$$6.5 \times 10^3 \cancel{KJ} \times \frac{2 \cancel{\text{mol}}}{571.6 \cancel{KJ}} \times \frac{18.0 g}{1 \cancel{\text{mol}}}$$

KJ - mol - gram

$$1 \times 10^5 \cancel{KJ} C_6H_6 \times \frac{1 \text{ mol}}{49.03 \cancel{KJ}} \times \frac{78.11 \text{ g}}{1 \text{ mol}}$$

C 12.01(6)
H 1.01(6)

Homework - Worksheet #1,2

Worksheet #3-6

What we've looked at so far...

- Energy changes when the temperature changes
  (heating water from 20 °C to 50°C)
- Energy changes when the temperature remains the same.
  (melting of ice at 0°C)

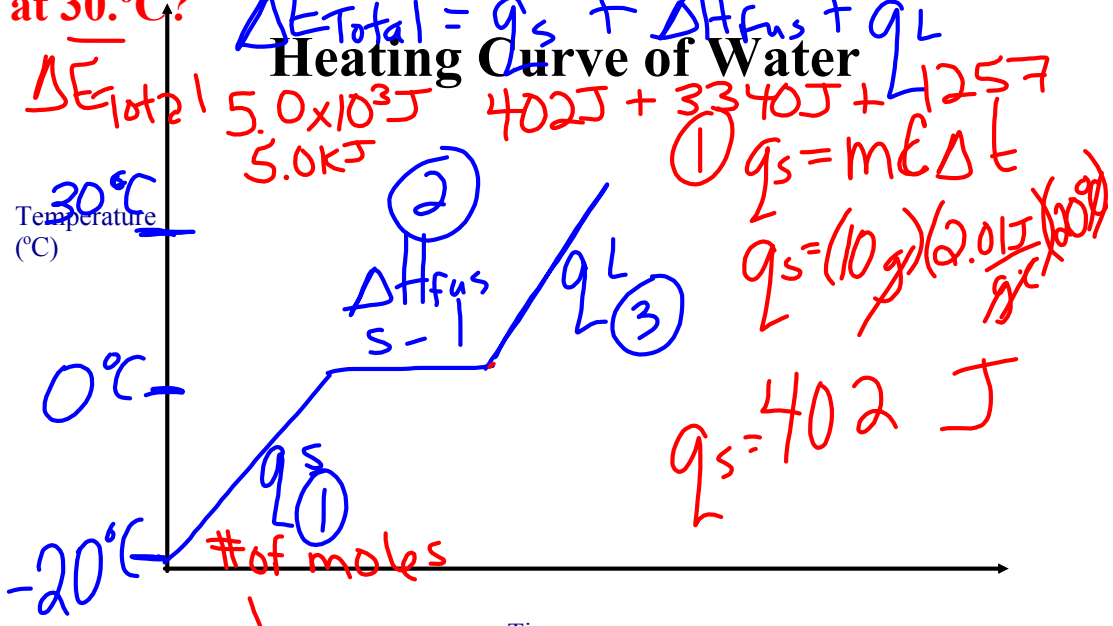
What if you heat 10. g of ice at -20. °C until it is water at 30.°C?

$\Delta E_{Total} = q_s + \Delta H_{fus} + q_L$

$\Delta E_{total} = 5.0 \times 10^3 J$

$5.0 kJ$

$402 J + 3340 J + 1257$



① $q_s = mc\Delta t$

$q_s = (10 g)(2.01 J/g°C)(20°C)$

$q_s = 402 J$

② $\Delta H = nH_{fus}$

$10 g \times \frac{1 mol}{18.02 g} \times (6.01 kJ/mol)$

$\Delta H_{fus} = 3.34 kJ \times \frac{1000 J}{1 kJ} = 3340 J$

③ $q_L = mc\Delta t$

$= (10)(4.19)(30)$

$q_L = 1257 J$

What we've looked at so far...

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  (heating water from 20 °C to 50°C)
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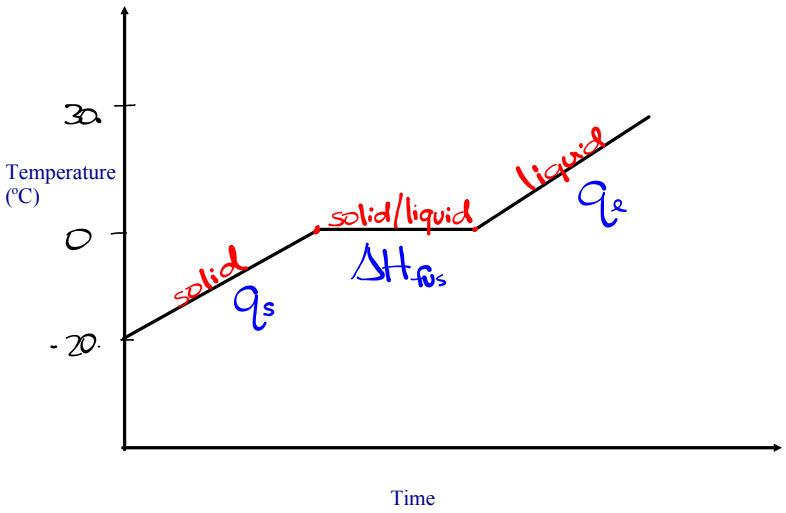
  (melting of ice at 0°C)

What if you heat 10. g of ice at -20. °C until it is water at 30.°C?

Heating Curve of Water



$\Delta E_T = q_s + \Delta H_{fus} + q_l$

$\Delta E_T = (0.402\,kJ) + (3.346\,kJ) + (1.257\,kJ)$

$\boxed{\Delta E_T = 5.0\,kJ}$

$q_s = mC\Delta T$

$q_s = (10.g)\left(2.01\frac{J}{g\cdot°C}\right)(20.°C)$

$q_s = 402\,J$

$\Delta H_{fus} = nH_{fus}$

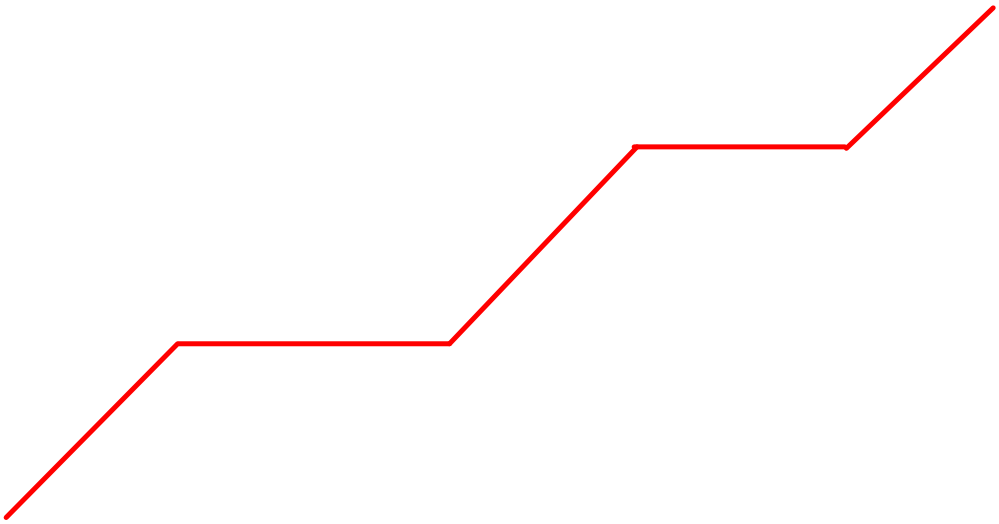
$\Delta H_{fus} = \left(\frac{10.g}{18.02\,g/mol}\right)\left(6.03\frac{kJ}{mol}\right)$

$\Delta H_{fus} = 3.346\,kJ$

$q_l = mC\Delta T$

$q_l = (10.g)\left(4.19\frac{J}{g\cdot°C}\right)(30.°C)$

$q_l = 1257\,J$



Total Energy Changes

Ex. Calculate the total energy change if 2.50 g of steam at 120.0°C is completely converted to ice at 0.0°C.

$q_g + \Delta H_{cond} + q_L + \Delta H_{solid}$

$\Delta E_{total} = -100.5J + -5646J + -1047.5J + -833J$

$-7.6 \times 10^3 J$

$7600 \div 1000$

① $q_g = mC\Delta t$

$(2.50)(2.01)(-20)$

$q = 100.5J$

② $\Delta H_{cond} = nH_{cond}$

# of moles

$2.50g\ H_2O \times \frac{1\ mol}{18.02g}$

$n = 0.139 mol$

$(0.139)(-40.7)$ x1000

$\Delta H_{cond} = -5.65 KJ$

$= -5650 J$

③ $q_L = mC\Delta t$

$(250)(4.19)(-100)$

$q = -1047.5J$

④ $\Delta H_{solid} = nH_{solid}$

$(0.139)(-6.01)$ x100

$\Delta H_{solid} = -0.833 KJ$

$\Delta H_{solid} = 833J$

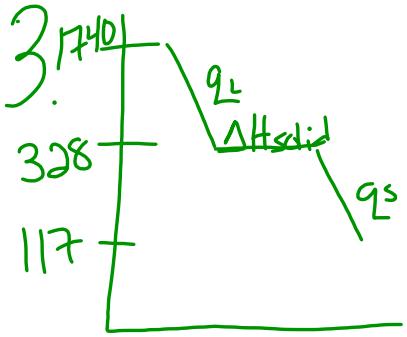
1. 200KJ
200000J

2. -31000 kJ
-31000000 J

3. -42.2 KJ
-42200 J

4. 4.00×10^2 KJ
4.00×10^5 J

$H_{solid\ Pb} = 4.77$ KJ/mol



Melting point 328

boiling point 1740

$\Delta E_{Total} = q_L + \Delta H_{solid} + q_s$

$q_L = mC\Delta t$

$(150g)(0.159 J/g°C)(-1412°C)$

$= -33676.2J / 1000$

$= -33.68KJ$

$\Delta H_{solid} = n H_{solid}$

$150g\ Pb \times \frac{1\ mol}{207.2g} = 0.7239\ mol$

$(0.7239 mol)(-4.77 KJ/mol)$

$= -3.453KJ$

$q_s = mC\Delta t$ (1)

$(150g)(0.159 J/g°C)(-211°C)$

$= -5032.35J$

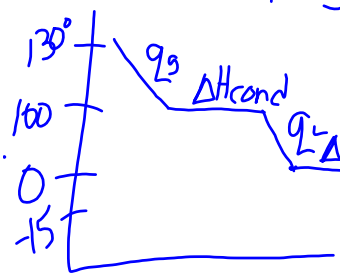
$= -5.03KJ$

$\Delta E_{Total} = (-33.68KJ) + (-3.453KJ) + (-5.03KJ)$

$= -42.2KJ$

2. $10\,Kg \times \frac{1000\,g}{1\,Kg} = 10000\,g$

$\Delta E_{Total} = q_g + \Delta H_{cond} + q_L + \Delta H_{solid} + q_s$



$q_g = mC\Delta t$

$(10000\,g)(2.01\,J/g\cdot°C)(-30°C)$

$= -603\,KJ$

$-603000\,J$

$\frac{10000\,g}{18.02\,g/mol} = n = 554.9\,mol$

$\Delta H_{cond} = nH_{cond}$

$(554.9\,mol)(-40.7\,KJ/mol)$

$= -22586\,KJ$

$q_L = mC\Delta t$

$(10000\,g)(4.195\,J/g\cdot°C)(-100°C)$

$= 4190\,KJ$

$4190000\,J$

$\Delta H_{solid} = nH_{solid}$

$(554.9\,mol)(-6.01\,KJ/mol)$

$= -3335.2\,KJ$

$q_s = mC\Delta t$

$(10000\,g)(2.01\,J/g\cdot°C)(-15°)$

$= -301.5\,KJ$

$= -301500\,J$

Worksheet

Total Energy Changes

Worksheet 55

1) 1273 KJ 1270 KJ

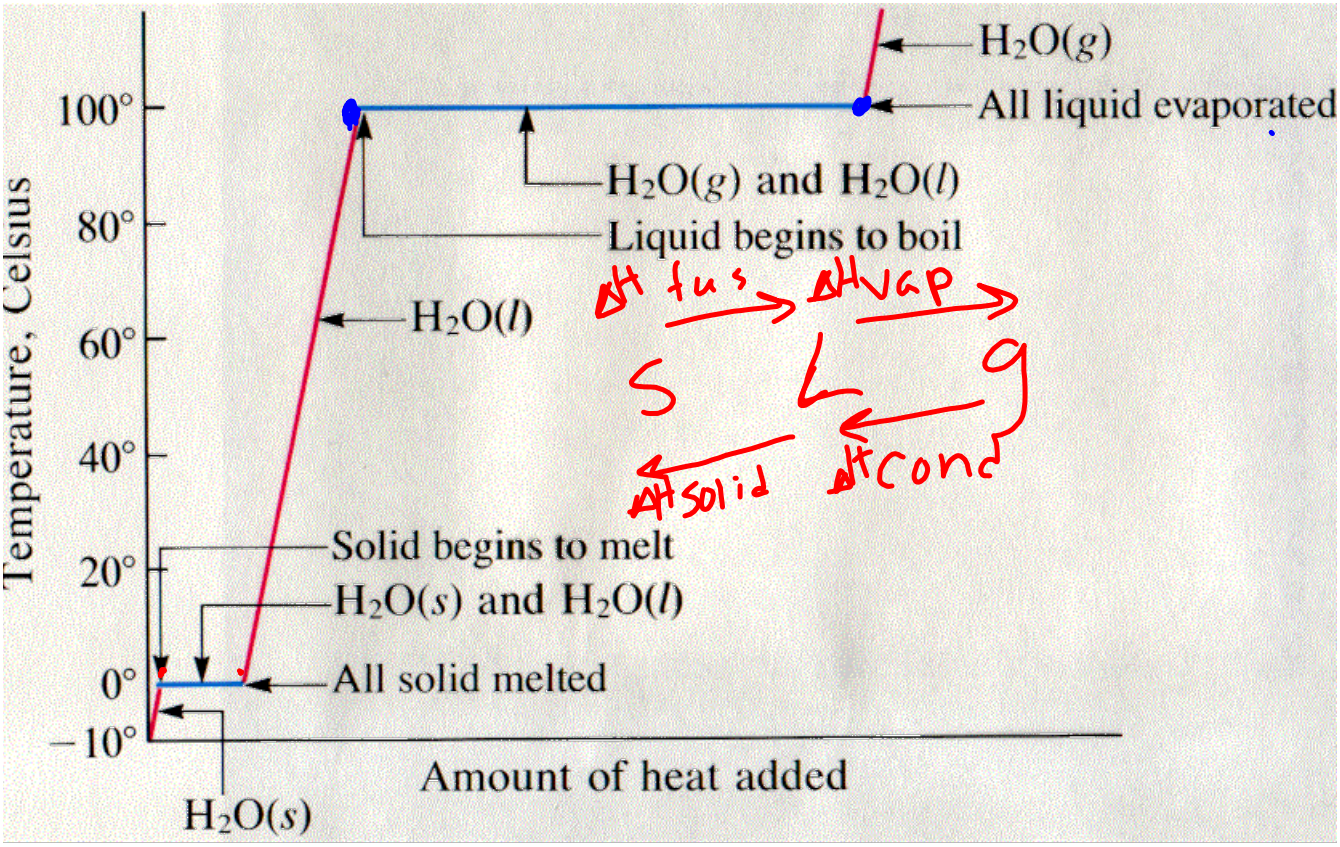
2) 7231 J 7000 J 7 KJ

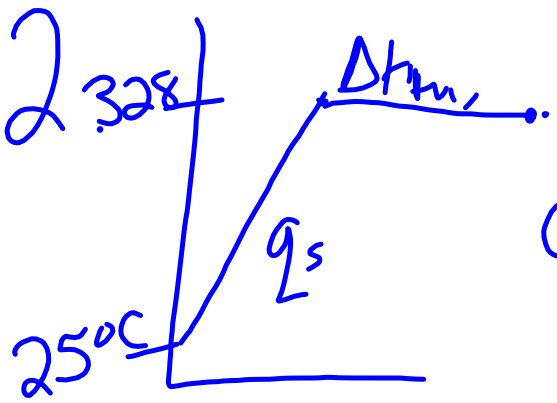
3) 478 KJ

4) -252670 KJ -300000 KJ

5) 539880 KJ 500000 KJ

Inert -A substance that is not chemically reactive.


$H_2O(g)$
All liquid evaporated
100°
80°
60°
40°
20°
0°
−10°
Temperature, Celsius
$H_2O(g)$ and $H_2O(l)$
Liquid begins to boil
$H_2O(l)$
ΔH fus
ΔHvap
S
L
g
ΔHsolid
ΔHcond
Solid begins to melt
$H_2O(s)$ and $H_2O(l)$
All solid melted
$H_2O(s)$
Amount of heat added



$\Delta E_{Total} = q_s + \Delta H_{fus}$

$(4818J) + (2413J)$

$7231J$

$q_s = mC\Delta t$

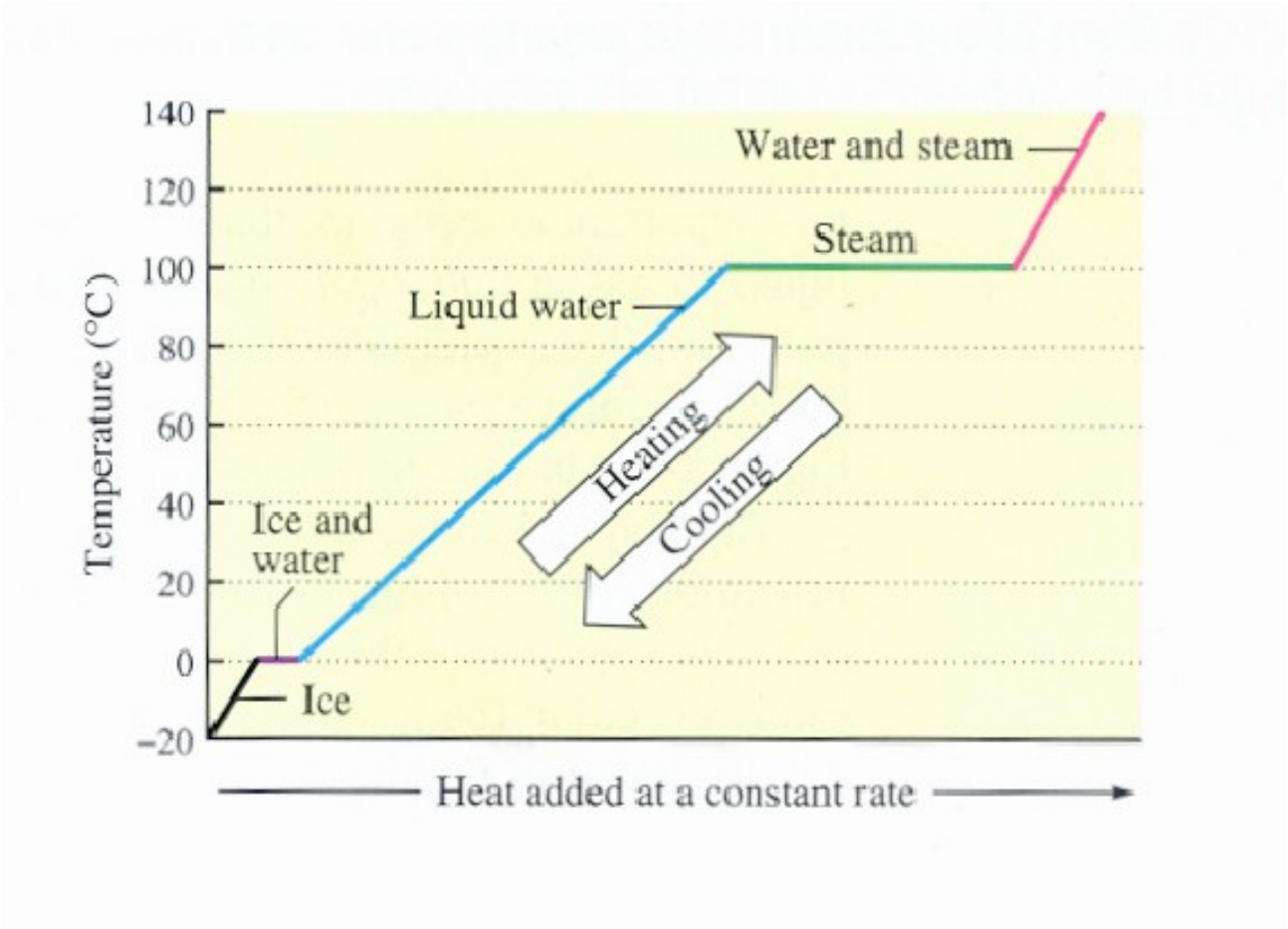
$(100g)\left(0.159\frac{J}{g°C}\right)(303°C)$

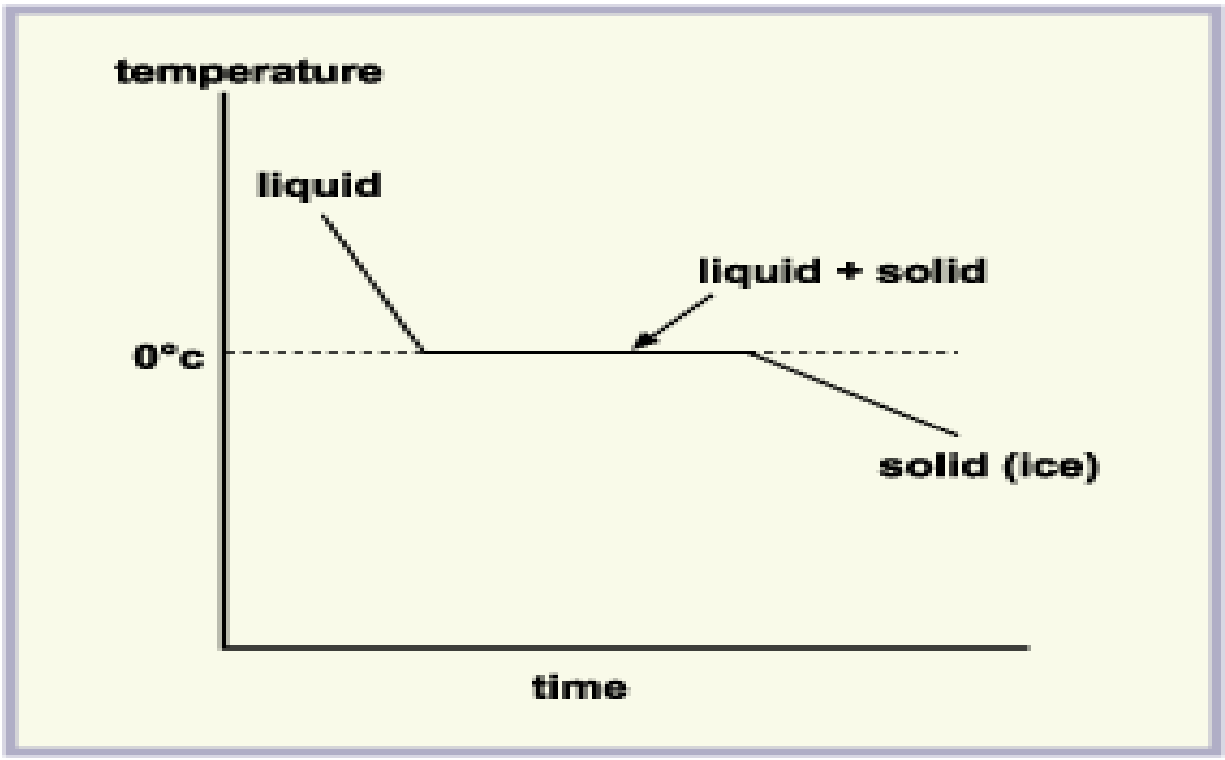
$= 4817.7J$

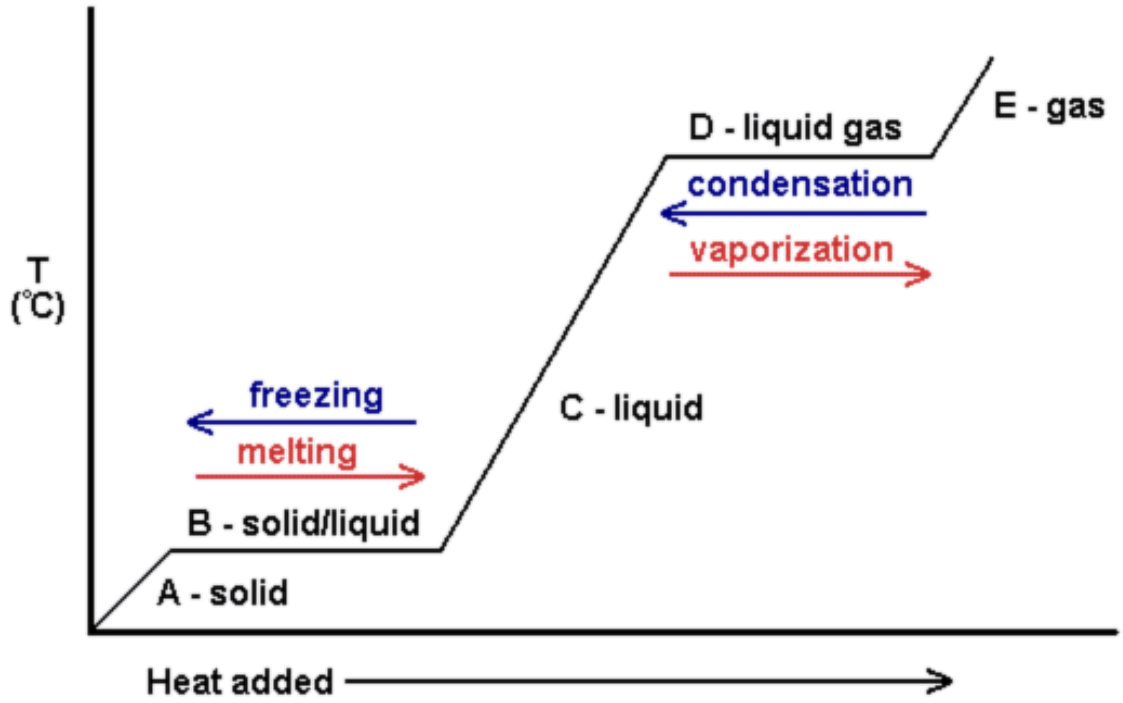
$\Delta H_{fus} = n H_{fus}$

$\left(5.0\frac{KJ}{mol}\right) 100g\,Pb \times \frac{1mol\,Pb}{207.2g\,Pb}$

$= 2.413KJ \times \frac{1000J}{1KJ} = 2413J$


140
120
100
80
60
40
20
0
−20
Temperature (°C)
Water and steam
Steam
Liquid water
Heating
Cooling
Ice and water
Ice
Heat added at a constant rate


temperature
liquid
liquid + solid
0°c
solid (ice)
time


T
(°C)
A - solid
B - solid/liquid
freezing
melting
C - liquid
D - liquid gas
condensation
vaporization
E - gas
Heat added