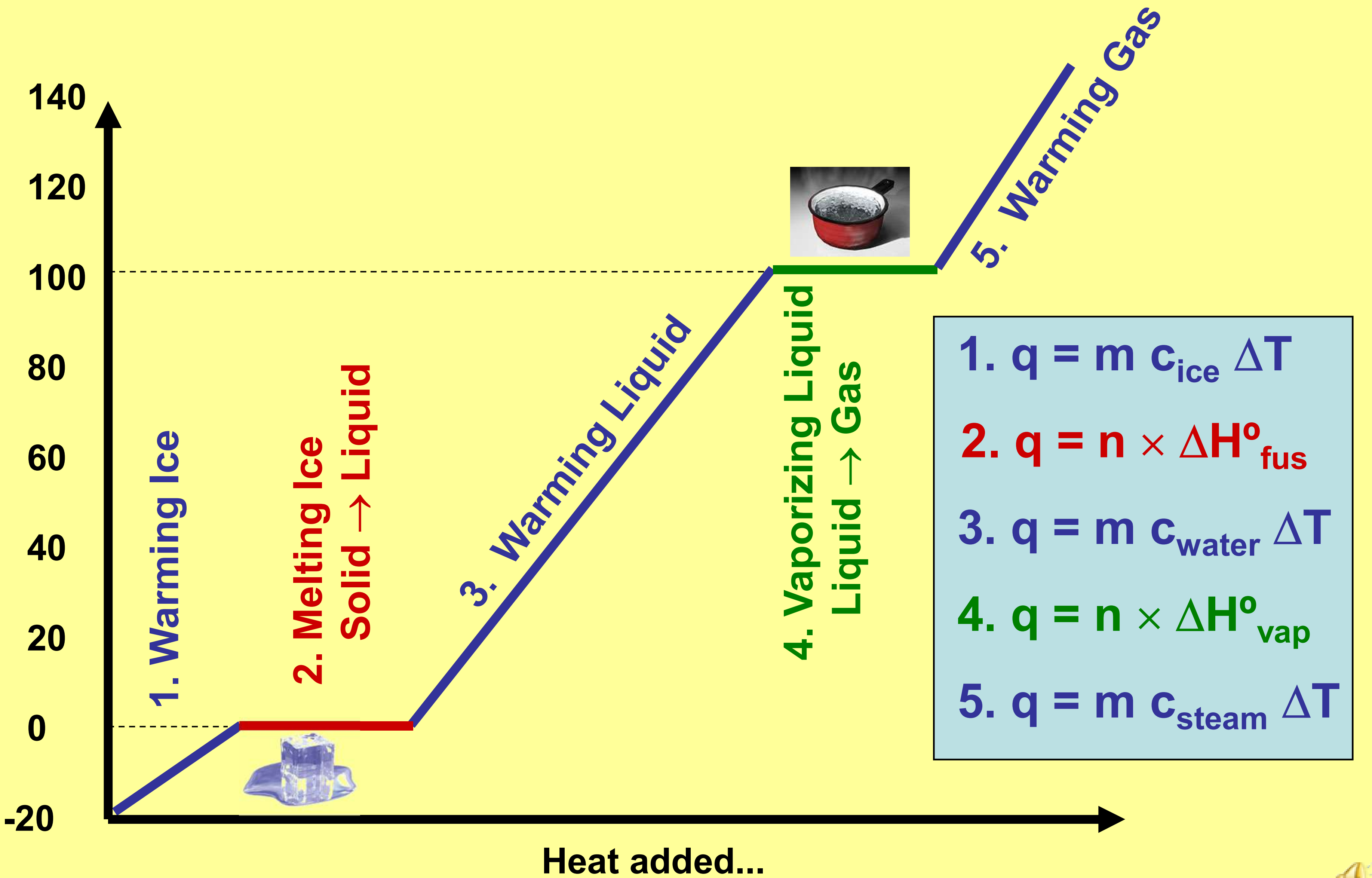
Phase Changes: Temperature Profile



Water: Thermodynamic Constants

$c_{ice} = 2.01 \text{ J/g}^oC$

$c_{water} = 4.184 \text{ J/g}^oC$

$c_{steam} = 2.16 \text{ J/g}^oC$

$\Delta H^o_{fus} = 6.02 \text{ kJ/mol}$

$\Delta H^o_{vap} = 40.17 \text{ kJ/mol}$

$\Delta H^o_{fus} << \Delta H^o_{vap}$

Why??

Vaporization *completely breaks all of the intermolecular forces.*

How much heat energy is required to convert 500. g of ice @ -10.ºC to steam @ 100. ºC ?

$$\underset{-10.0\ ^\circ C}{H_2O_{(s)}} \xrightarrow{1.} \underset{0.0\ ^\circ C}{H_2O_{(s)}} \xrightarrow{2.} \underset{0.0\ ^\circ C}{H_2O_{(l)}} \xrightarrow{3.} \underset{100.0\ ^\circ C}{H_2O_{(l)}} \xrightarrow{4.} \underset{100.0\ ^\circ C}{H_2O_{(g)}}$$

1. $q = m\, c_{ice}\, \Delta T$
2. $q = n \times \Delta H^o_{fus}$
3. $q = m\, c_{water}\, \Delta T$
4. $q = n \times \Delta H^o_{vap}$
5. $q = m\, c_{steam}\, \Delta T$

$c_{ice} = 2.01\ J/g^oC$

$\Delta H^o_{fus} = 6.02\ kJ/mol$

$c_{water} = 4.184\ J/g^oC$

$\Delta H^o_{vap} = 40.17\ kJ/mol$

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1. $q = m\ c_{ice}\ \Delta T$ $= 500.g \times 2.10\ J/g^\circ C \times (0 - -10.0^\circ C) =$ 10500 J

2. $q = n \times \Delta H^\circ_{fus}$ $= 27.754\ mol \times 6.02\ kJ/mol$ $=$ 167.08 kJ

3. $q = m\ c_{water}\ \Delta T$ $= 500.g \times 4.184\ J/g^\circ C \times (100 - 0^\circ C) =$ 209200 J

4. $q = n \times \Delta H^\circ_{vap}$ $= 27.754\ mol \times 40.17\ kJ/mol$ $=$ 1114.88 kJ

5. $q = m\ c_{steam}\ \Delta T$ *Not used in this problem.*

1.50×10^3 kJ