Chemistry *Estimated time for completion:* *60 min*

Name: Date:

§ 16.etc

Thermochemistry: Heating & Temperature

Most Important Ideas:

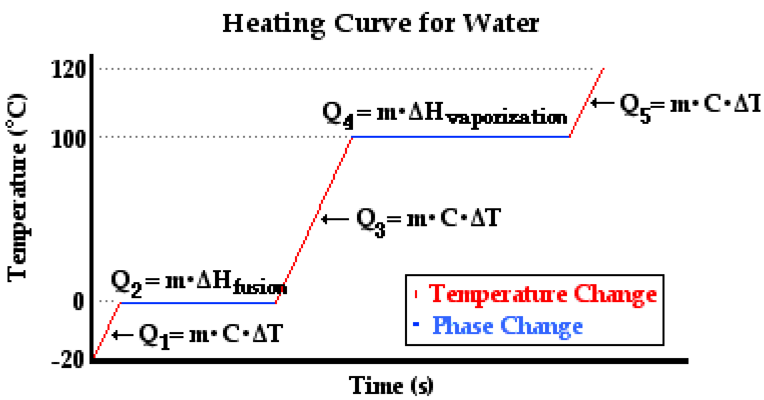
Objective

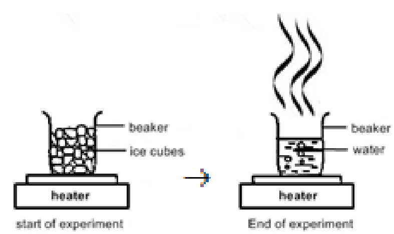
The objective of this activity is to understand the difference between heat and temperature. Specifically, you will calculate the heat required to amount of heat required for the phase changes associated with changing ice to steam.

Background

One common misconception is that heat and temperature is the same. They are related but they are not synonymous. Temperature is the measure of the average amount of kinetic energy in a substance. Heat affects temperature by speeding up, or slowing down, the average amount of kinetic energy. But heat cannot be measured directly because it includes potential energy (e.g., the energy stored in chemical bonds.

When the heat (q) does not involve a change in change pressure or temperature, it is called enthalpy and is symbolized as ΔH. ΔH and q are almost interchangeable. In this experiment, we will follow the heat (ΔH) involved in heating ice to steam. The curve of which is shown below.





Q_1 = heating ice from –20^{o}C to 0^{o}C.

Q_2 = melting ice to water at 0^{o}C.

(N.B., heat is still added but there is no change in temperature.)

Q_3 = heating water from 0^{o}C to 100^{o}C.

Q_4 = evaporating water to steam at 100^{o}C.

(N.B., again, heat is still added but there is no change in temperature.)

Q_5 = heating steam from 100^{o}C to120^{o}C

N.B. Always draw a picture when you can! And use it to organize your problem solving.

Values You'll Need to Solve These Problems

Specific Heats for the States of Water

ice	(s_{ice})	2.20 J/g ^oC
water	(s_{ice})	4.184 J/g ^oC
steam	(s_{steam})	2.01 J/g ^oC

Latent Heats of Fusion and Vaporization for Water:

Term	Transition	Value (J/g)
heat of fusion (ΔH_{fus})	ice → water	334
	water → ice	–334
heat of vaporization (ΔH_{vap})	water → steam	2260
	steam → water	–2260

Model - 1: Heating Ice at –20^{o}C to Steam at 120^{o}C

Problem

How much heat (kJ) is required to heat a 100.-g sample of ice to steam at 120^{o}C?

***Solution** – Refer back to the heating curve to solve*

1. Draw and label the heating curve, as is shown in the above figure.
2. Calculate one step at a time.

Q_1 ice @ –20 ^{o}C → 0^{o}C

$$= m\ s\ \Delta T = (100g)\left(2.20\frac{J}{g\ ^\circ C}\right)(0-(-20)\ ^\circ C) = 4{,}400\ J$$

Q_2 ice @ 0 ^{o}C → water @ 0^{o}C

There is no temperature change because all of the heat went into breaking the hydrogen–bonds that turned the ice into liquid water. The heat associated with this is called the 'latent heat of fusion' or simply, the 'heat of fusion'. Solve the phase changes by using dimensional analysis.

$$= m\ \Delta H_{fus} = (100.\ g)\left(334\frac{J}{g}\right) = 334{,}000\ J$$

Q_3 water @ 0 ^{o}C → 100^{o}C

$$= m\ s\ \Delta T = (100g)\left(4.184\frac{J}{g\ ^\circ C}\right)(100-0\ ^\circ C) = 41{,}840\ J$$

Q_4 water @ 100 ^{o}C → steam @ 100^{o}C

$$m\ \Delta H_{vap} = (100.\ g)\left(2260\frac{J}{g}\right) = 226{,}000\ J$$

Q_5 steam @ 100 ^{o}C → steam 120^{o}C

$$= m\ s\ \Delta T = (100g)\left(2.01\frac{J}{g\ ^\circ C}\right)(120-100\ ^\circ C) = 4{,}020\ J$$

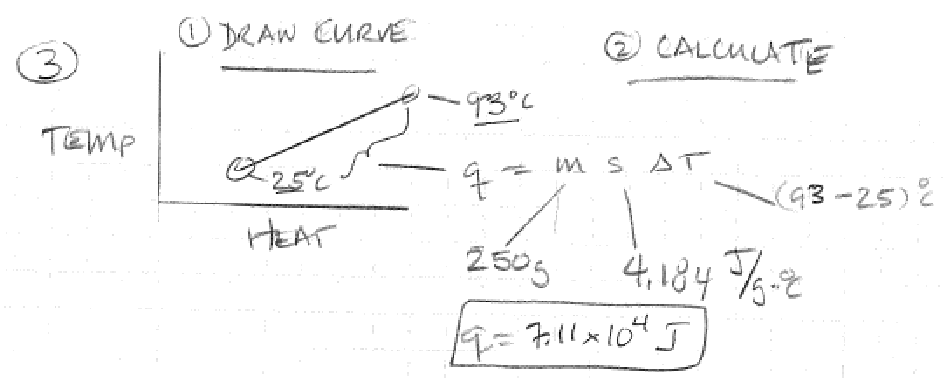
Total: 610,260 J $x\ \frac{1\ kJ}{1000\ J}$ = <u>610.3 kJ</u>

Problems Use the values in the data table above for specific heats and latent heats.

1. What is the highest latent heat? Why is it the highest?
The quantity of **heat** absorbed or released by a substance undergoing a change of state, such as ice changing to water or water to steam, at constant temperature. Is associated with the breaking and forming of bonds (for water: H-bonds).

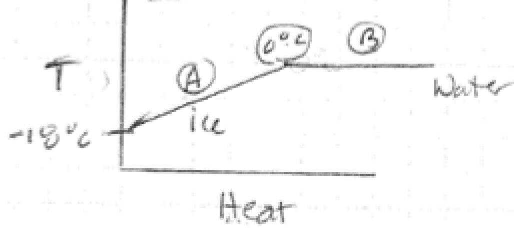
2. How much heat (J) does it take to raise temperature of one gram of water by 1^oC?
By definition, one calorie. One cal = 4.184 J.

3. A cup of Starbuck's coffee is served at about 93 ^oC. How much heat (J) is required to raise the temperature of a 250-mL cup of coffee starting at 25^oC room temperature? (Assume coffee is water.)



4. Summer is coming and, with it, picnics. You buy a 10-lb bag of ice at the convenience store. The temperature of the ice is –18^{o}C. How much heat (kJ) is given off it melts to 0^{o}C? (2.20 lb = 1 kg)

(4) (1) CURVE



(1) $m = \frac{10 lb}{1} \times \frac{1 kg}{2.20 lb} \times \frac{10^3 g}{1 kg} = 4.546 \times 10^3 g$

(A) ice — $-18°C \rightarrow 0°$

$q = m s \Delta T = (4.55 \times 10^3)(2.20 \frac{J}{g°C})(18°C) = 180,180 J$

(B) ice → water

$q = (4.546 \times 10^3 g)(\frac{334 J}{g°C}) \Rightarrow$ = 1,518,364 J

1,698,544 J

= **1,699 kJ**

5. When there is overnight frost coming to an orange orchard, it is not uncommon to spray the fruit with water.
   a. To answer this, calculate the heat given off by the freezing (0^oC) of 10 mL of water.
   b. Why do pomologists spray the fruit to keep it from freezing?
   c. What was the molecular source of the released heat?

(5) water $\xrightarrow{0°C}$ ice

(A)

⇒ no ΔT

$q = \left(10.0\,mL \times \frac{1\,g}{1\,mL}\right)\left(334\,\frac{J}{g}\right)$

(for H_2O)

$q = 3340\,J$

(B) Heat is given off by the water freezing to counteract the freezing temperature outside.

(C) Heat is added to ice to break apart some of the H-bonds when it melts to form water.

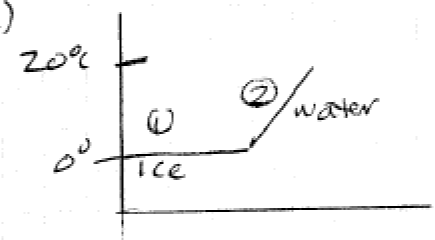
When the water freezes, this heat is released when H-bonds form making ice.

[Law of conservation of energy]

6. A 25.00 gram sample of ice at 0.0°C melts and then warms up to 20.0°C.
   a. Draw the heating curve and label different stages. Add the formula for each to be used in the calculations.
   b. How much energy is absorbed?

⑥ 25.0 g ice @ 0°C → water → 20°C

(a)



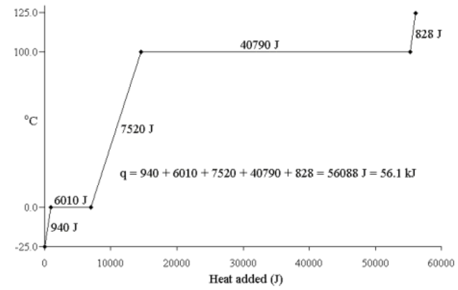
(b) ① $q = (25.0\,g)\left(334\,\frac{J}{g}\right) = \underline{8{,}350\,J}$

② $q = m\,s\,\Delta T = (25.0\,g)\left(4.184\,\frac{J}{g\,°C}\right)(20\,°C)$

$q = \underline{2{,}092\,J}$

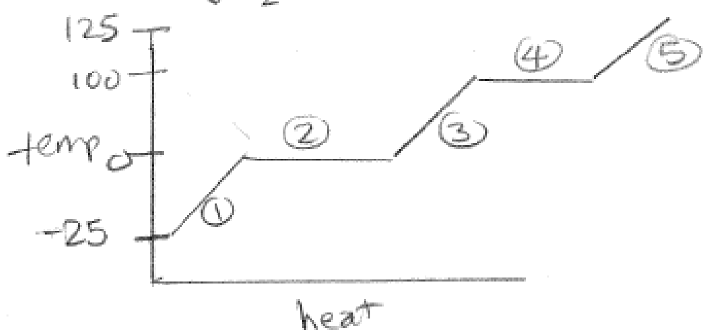
$\boxed{\Sigma = 10{,}442\,J}$

7. How many joules of heat are required to heat 18.0 g of water from –25°C to 125°C?



(7)

18.0 g H_2O: -25°C to 125°C



(1) ice -25 to 0°C ⇒ $msΔT = (18.0g)(2.20 \frac{J}{g°C})(25°C) = 990.0 J$

(2) ice → water ⇒ $q = (18.0g)(334 \frac{J}{g}) = 6,012.0 J$

(3) water 0→100°C ⇒ $msΔT = (18.0g)(4.184 \frac{J}{g°C})(100°C) = 7,531.2 J$

(4) water → steam ⇒ $q = (18.0g)(2260 \frac{J}{g}) = 40,680 J$

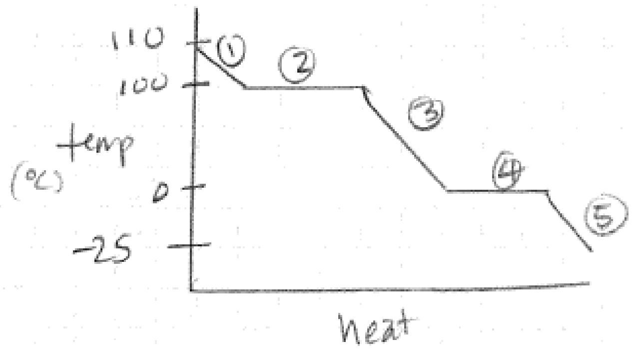
(5) steam 100→125°C ⇒ $q = msΔT = (18.0g)(2.01 \frac{J}{g°C})(25°C) = 904.5 J$

TOTAL = 56,117.7 J

$q = 56.12 kJ$

8. How many joules of heat are given off when 1 ton (2,000 lb; 1 lb = 454 g) of steam at 110°C are cooled to –25°C?

(8) 1 ton steam (110°) → ice (−25°C)



MASS: $\frac{1 \text{ ton}}{1} \times \frac{1{,}000 \text{ lb}}{1 \text{ ton}} \times \frac{454 \text{ g}}{1 \text{ lb}} = 4.54 \times 10^5 \text{ g}$

(1) Steam 110 → 100°C

$q = ms\Delta T = \left(\frac{4.54 \times 10^5 \text{ g}}{1}\right)\left(\frac{2.01 \text{ J}}{\text{g °C}}\right)(10°\text{C}) = 9.125 \times 10^6 \text{ J}$

(2) steam → water $q = (4.54 \times 10^5 \text{ g})(2260 \text{ J/g}) = 1.026 \times 10^9 \text{ J}$

(3) water 100° → 0°C

$q = ms\Delta T = (4.54 \times 10^5 \text{ g})(4.184 \frac{\text{J}}{\text{g °C}})(100°) = 1.900 \times 10^8 \text{ J}$

(4) water → ice $q = (4.54 \times 10^5 \text{ g})(334 \frac{\text{J}}{\text{g}}) = 1.516 \times 10^8 \text{ J}$

(5) ice 0° → −25°C

$q = ms\Delta T = (4.54 \times 10^5 \text{ g})(2.20 \frac{\text{J}}{\text{g}}) = \quad = 9.988 \times 10^5 \text{ J}$

$\Sigma = 1.378 \times 10^9 \text{ J}$

$= 1.378 \times 10^6 \text{ kJ}$