Key Questions:

1. What does it mean to say that a calorimeter is a "closed system?"

all the heat lost by the reaction is gained by the water

2. Examine the calorimeter, if the chemical reaction run in the reaction vessel is endothermic, what will happen to the temperature of the surrounding water? Why?

it will decrease because heat is being taken from the water to be used in the reaction

3. Examine the calorimeter, if the chemical reaction run in the reaction vessel is exothermic, what will happen to the temperature of the water? Why?

it will increase because heat is being released by the reaction so the water absorbs it

4. What is the numerical value and units of the specific heat capacity of water? (look in reference packet)

4.18 J/g°C

5. What information does the specific heat capacity of water provide? (try to figure out what the units are telling us)

the amount of energy needed to raise 1 gram of a substance 1°C

6. How can the heat released into some mass of water be calculated from the specific heat capacity of water and the change in temperature of the water? (Answer in words not with an equation.)

all the heat transfered from the system and surroundings has to be equal

7. A calorimeter was used to measure the heat released by a chemical change. The calorimeter contained 100.00 g of water at an initial temperature of 10.0 0C. When the reaction was finished the temperature of the water increased to 75.0 0C. What is the heat quantity released by the chemical change? (Include correct unit)

$q = mC\Delta T$

$q = (100.00g)(4.18\ J/g\ °C)(75.00°C - 10.00°C)$

$q = 27170\ J$

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8. If a substance with a larger specific heat than water were used in the experiment, identify whether ΔT would be larger or smaller. Explain.

the ΔT would be smaller because it would require more energy to raise the temp 1°C

Problem

1. Calculate the change in temperature that results from the addition of 2500 J of heat energy to a 25-g sample of water.

$q = 2500\,J$
$m = 25\,g$
$c = 4.18\,J/g°C$
$\Delta T = $

$q = mc\Delta T$
$(2500) = (25)(4.18)(\Delta T)$
$\Delta T = 23.9\,°C$

(b) Calculate the change in temperature that results from the addition of 2500 J of heat energy to a 25-g sample of lead, Cp = 0.130 J/g°C.

$q = 2500\,J$
$m = 25\,g$
$c = 0.130\,J/g°C$
$\Delta T = ?$

$q = mc\Delta T$
$(2500\,J)\ (25\,g)(0.130\,J/g°C)(\Delta T)$
$\Delta T = 769\,°C$

(c) Explain the difference in the temperature changes, found in parts (a) and (b), in terms of the specific heat capacities of lead and water.

- lead had a much lower specific heat so the change in temp was much higher
- water has a much higher specific heat, so the change in temp is much less

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CW: Quantities of Heat

Specific Heat

1. If 25g of liquid water is heated from 10.0°C to 25.0°C, what is the heat change in Joules? Is this an endothermic change or an exothermic change? (1600 J, endothermic)

$q = ?$
$m = 25\ g$
$c = 4.18\ J/g^\circ C$
$\Delta T = 25.0^\circ C - 10.0^\circ C = 15.0^\circ C$

$q = (25g)(4.18\ J/g^\circ C)(15.0^\circ C)$
$q = 1600\ J$

2. What mass of water can be heated from 45.0°C to 70.0°C by the addition of 875 calories? (35g)

need to convert to Joules

$q = 875\ \text{calories} \times \frac{4.184\ J}{1\ \text{calorie}} = 3661\ J$
$m = ?$
$c = 4.18\ J/g^\circ C$
$\Delta T = 70.0^\circ C - 45.0^\circ C = 25.0^\circ C$

$3661\ J = (m)(4.18\ J/g^\circ C)(25.0^\circ C)$
$m = 35.0g$

3. If 700.0 g of water at 90.0°C loses 27 kJ of heat energy, what is its final temperature? (81°C)

$q = 27\ kJ \times \frac{1000\ J}{1\ kJ} = 27000\ J$
$m = 700.0g$
$c = 4.18\ J/g^\circ C$
$\Delta T = x - 90.0^\circ C$

$-27000\ J = (700.0g)(4.18\ J/g^\circ C)(x - 90.0^\circ C)$
$-9.2 = x - 90.0^\circ C$
$x = 80.8^\circ C$

4. A quantity of water is heated from 10.0°C to 50.0°C. During the process, 50.0 kJ of heat energy is added to the water. How many grams of water are heated? (299.g)

$q = 50.0\ kJ \times \frac{1000\ J}{1\ kJ} = 50000\ J$
$m = ?$
$c = 4.18\ J/g^\circ C$
$\Delta T = 50.0^\circ C - 10.0^\circ C = 40.0^\circ C$

$50000\ J = m(4.18\ J/g^\circ C)(40.0^\circ C)$
$m = 299\ g$

5. What is the specific heat of an unknown substance if the addition of 950 J of heat energy caused a 20.0 g sample to warm form 18.0°C to 42.0°C? (2.0 J/g°C)

$q = 950\ J$
$m = 20.0g$
$c = ?$
$\Delta T = 42.0^\circ C - 18.0^\circ C = 24.0^\circ C$

$950\ J = (20.0g)(x)(24.0^\circ C)$
$x = 2.0\ J/g^\circ C$

6. Mercury has a specific heat of 0.14 J/g°C. How much energy is required to increase the temperature of a 22.8 g sample of mercury from 16.1°C to 32.5°C? (52 J)

$q = ?$
$m = 22.8g$
$c = 0.14\ J/g^\circ C$
$\Delta T = 32.5^\circ C - 16.1^\circ C = 16.4^\circ C$

$q = (22.8g)(0.14\ J/g^\circ C)(16.4^\circ C)$
$q = 52\ J$

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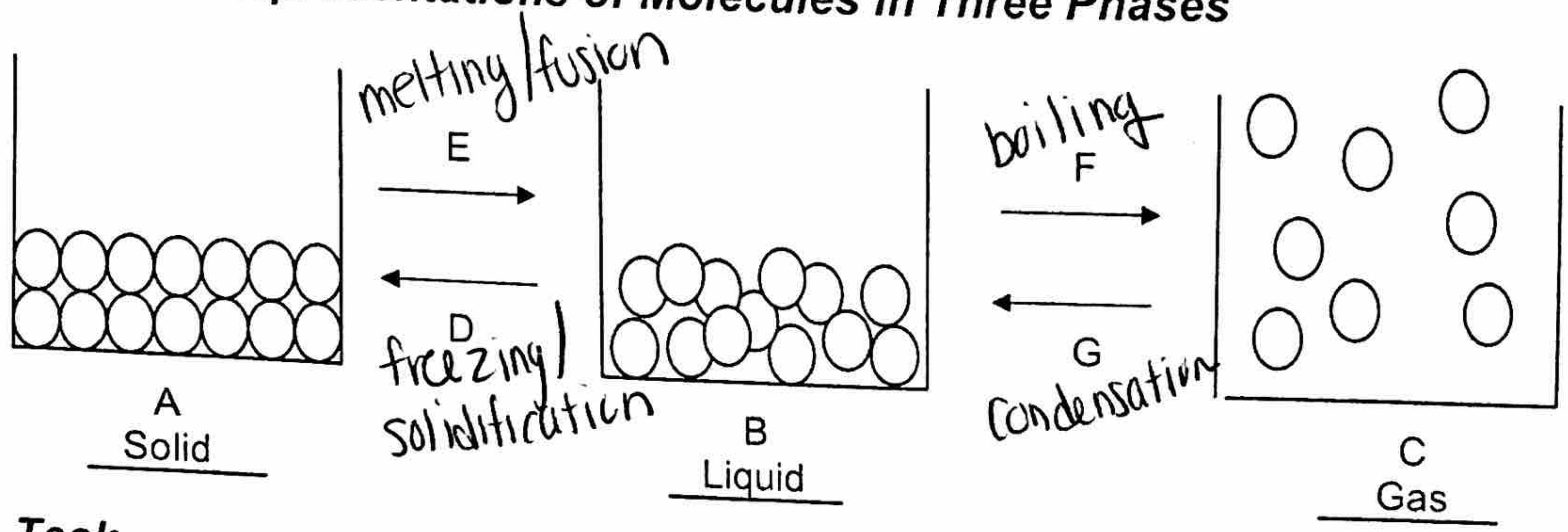
Model 1: Representations of Molecules in Three Phases



Task

Label each arrow (D, E, F, G) in Model 1 with the appropriate phase change (fusion/melting, solidification/freezing, boiling, condensation).

Key Questions

1. Which arrows in Model 1 indicate the addition of energy?

E, F

2. Which term, endothermic or exothermic, is used to describe the situation when energy is added into a system from the surroundings?

endothermic

3. Which arrows in Model 1 indicate the release of energy?

D, G

4. Which term, endothermic or exothermic, is used to describe the situation when energy is released into the surroundings by the system?

exothermic

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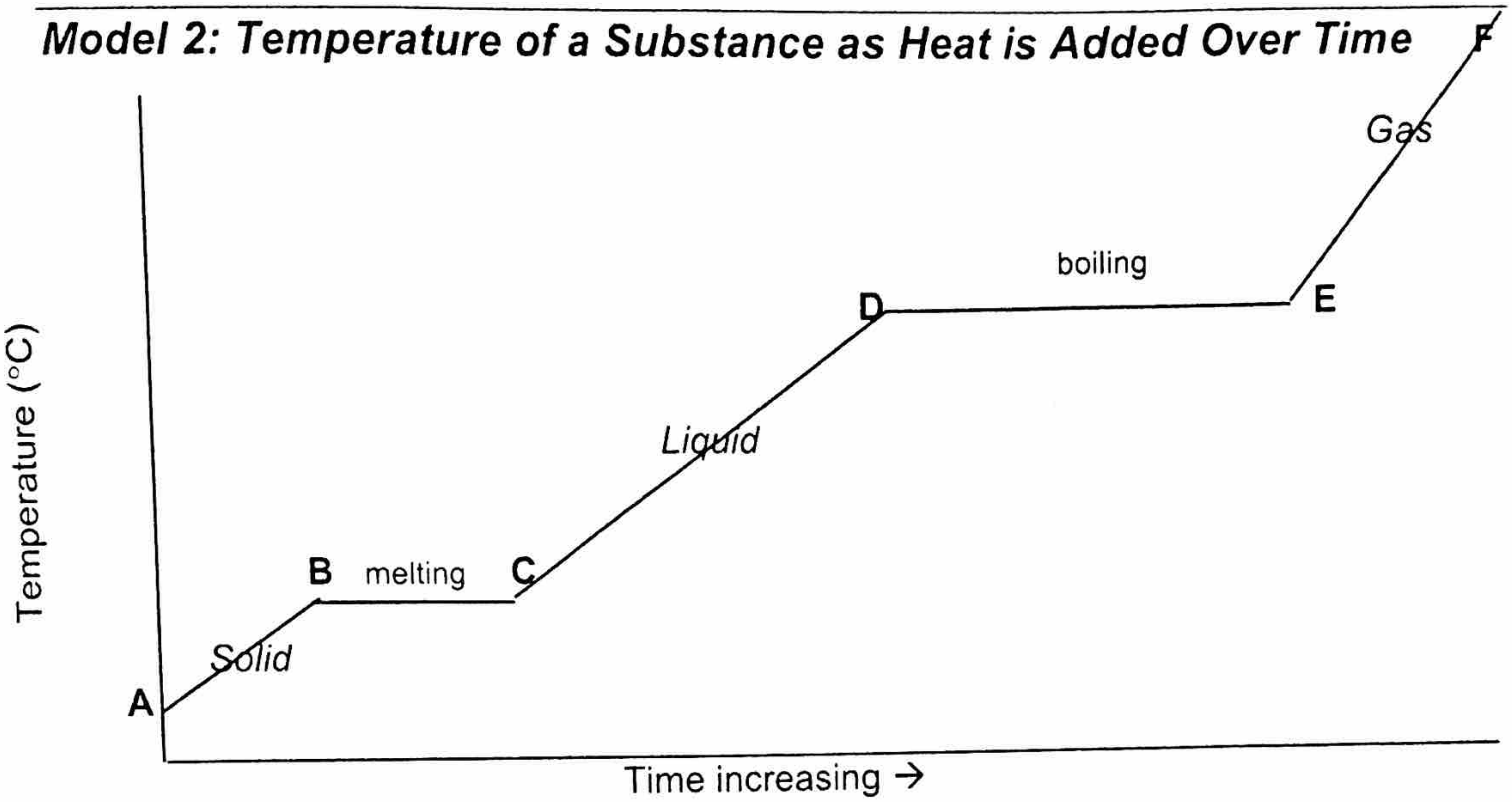
Model 2: Temperature of a Substance as Heat is Added Over Time



Key Questions

1. What is plotted on the x-axis and what is plotted on the y-axis of the graph in Model 2?

   x-axis (time) y-axis (temp)

2. During which line segments does temperature increase?

   AB, CD, EF

3. During which line segments is there no change in temperature?

   BC, DE

4. If this substance were water, at what temperature would segment B – C occur?

   0°C

5. If this substance were water, at what temperature would segment D – E occur?

   100°C

POGIL 2005, 2006
Authored by: Erin Graham; Revised by: Josephine Parlagreco, Lizabeth Tumminello
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6. On the molecular level, why is energy added in order to complete an endothermic phase change? (Refer to both Models in your answer.)

   molecules need to move faster, so they need more energy

7. On the molecular level, why is energy released to complete an exothermic phase change? (Refer to both Models in your answer.)

   molecules need to slow down, so they need to release energy

8. Comparing segments B – C and D – E, what information is conveyed by the observation that segment D – E is longer?

   it takes more energy to to convert from l → g than s → l

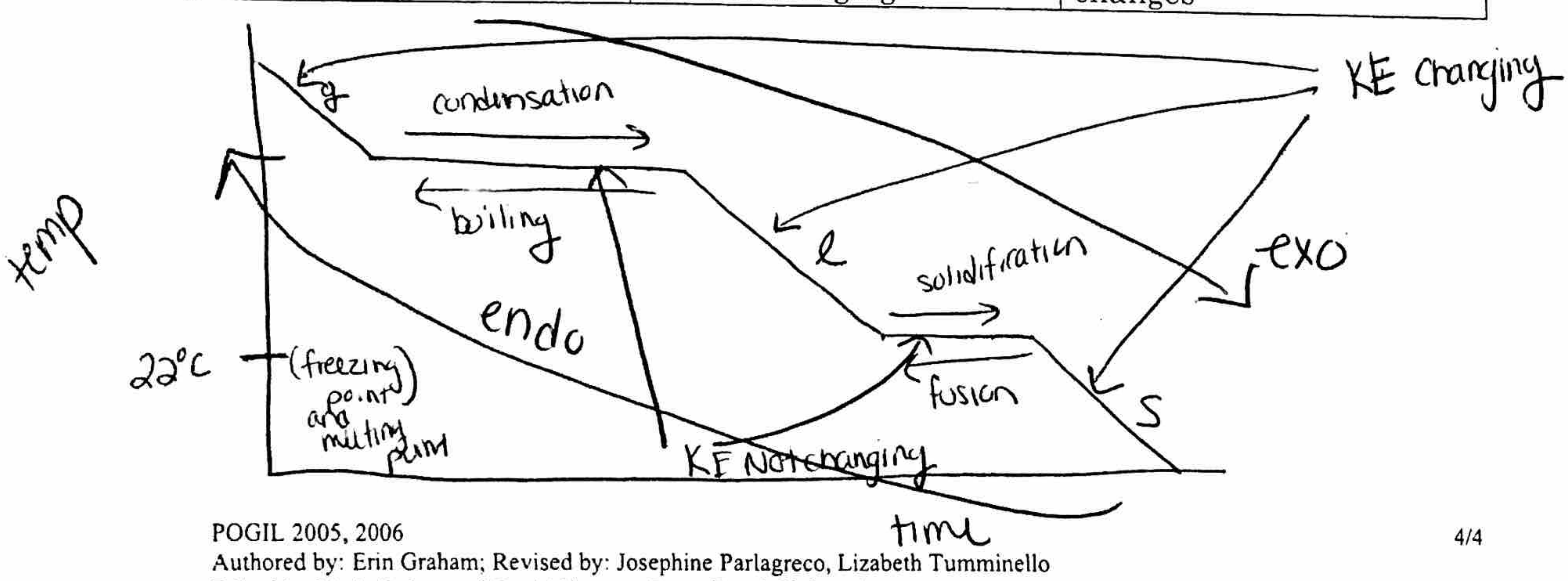
Exercise

A sample of a mythical substance is cooled from a temperature of 250°C to 10°C in two hours. The boiling point of the substance is 175°C and the melting point is 22°C.

Using this information, draw a cooling curve for the sample.

On the curve clearly label the following items in the appropriate locations (use arrows as needed to indicate direction or exact location on the curve. Some terms may be used more than once, as needed.):

~~Solid~~ ~~Liquid~~ ~~Gas~~ ~~Condensation~~ ~~Boiling~~	~~Freezing Point~~ ~~Fusion~~ ~~Solidification~~ ~~KE changing~~ ~~KE not changing~~	~~Melting Point~~ Direction of endothermic changes Direction of exothermic changes



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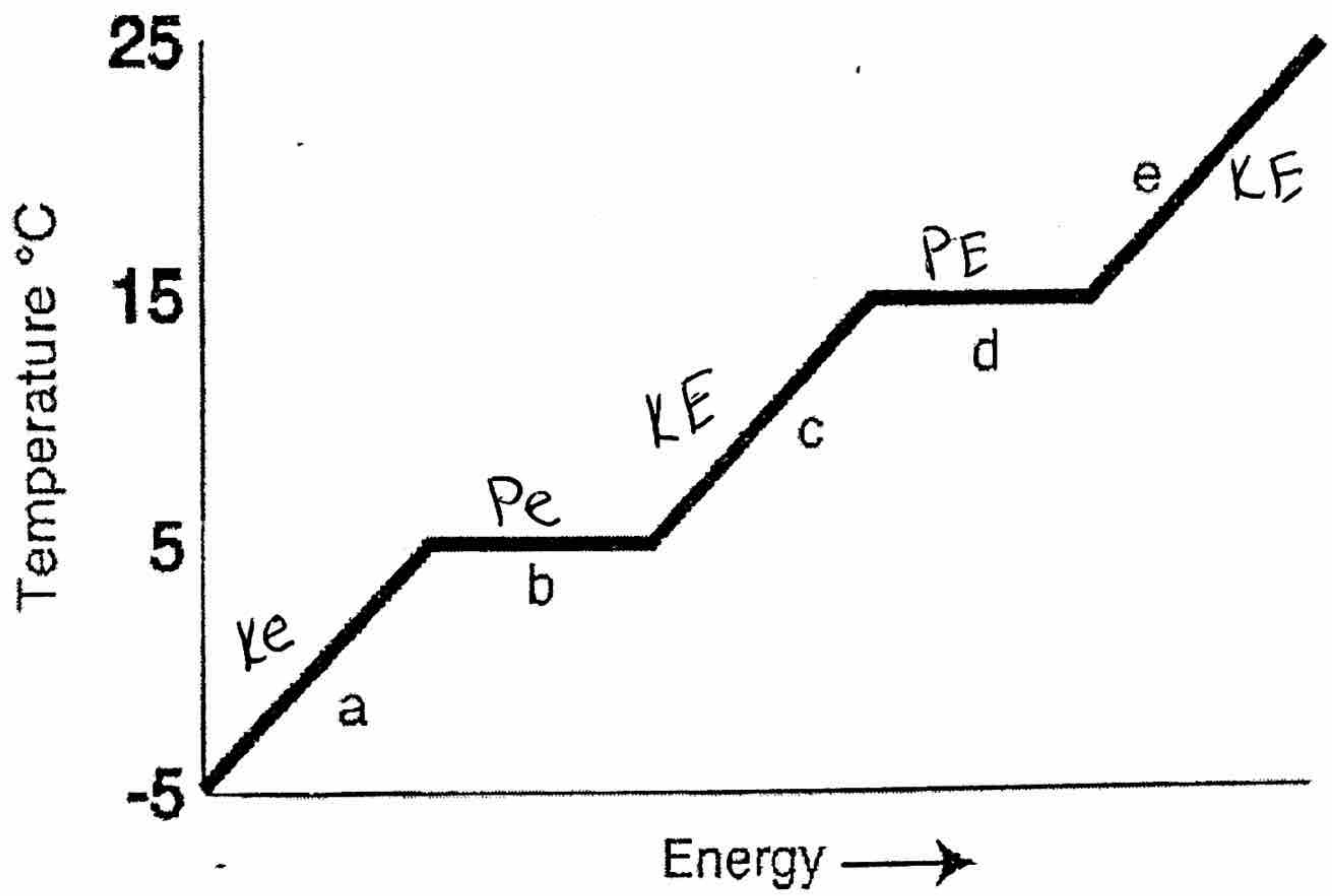
Heating Curve Homework

Name____________________

Date____________Per____



ANSWER THE FOLLOWING USING THE ABOVE HEATING CURVE

1. What is the melting temperature of the above substance? 5°C
2. What is the freezing temperature of the above substance? 5°C
3. What is the boiling temperature of the above substance? 15°C
4. The part of the graph labeled "e" represents temperatures at which gas is being heated. Describe what is happening for each of the other lettered sections of the graph:

a. Solid is being heated
b. Solid is melting to a liquid (IMF is breaking)
c. liquid is being heated
d. liquid is being evaporated to steam (IMF is breaking)

5. In which section of the graph are atoms moving the least? a
6. In which section of the graph is this substance all liquid? c
7. Write KE where kinetic energy is changing and "PE " where potential energy is changing.

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Phase Changes

1. What quantity of heat is required to vaporize 600. grams of water at 100°C? (1356 kJ)

$\Delta H = mH_v$

$\Delta H = (600.g)(2260 J/g)$

$\Delta H = 1356000 J$ or $1356 kJ$

$\rightarrow 1360 kJ$

2. A quantity of water vapor at 100°C is condensed to liquid water at 100°C. This process releases 6.50 x10^4 J of energy. What mass of water has been condensed? (28.8g)

$\Delta H = mH_v$

$-6.50 \times 10^4 J = x(-2260 J/g)$

$x = 28.8 g$

3. How much heat (in kJ) is released when 85.0 g of liquid water freezes to ice, at 0°C? (28.4 kJ)

$\Delta H = mH_f$

$\Delta H = (85.0g)(334 J/g)$

$\Delta H = 28390 J \rightarrow 28.4 kJ$

add a "0" after the 6.8

4. What quantity of ice, at 0°C, will be melted by 6.80 x 10^4 J of energy? (204g)

$\Delta H = mH_f$

$-6.8 \times 10^4 J = m(-334 J/g)$

$m = 204g$

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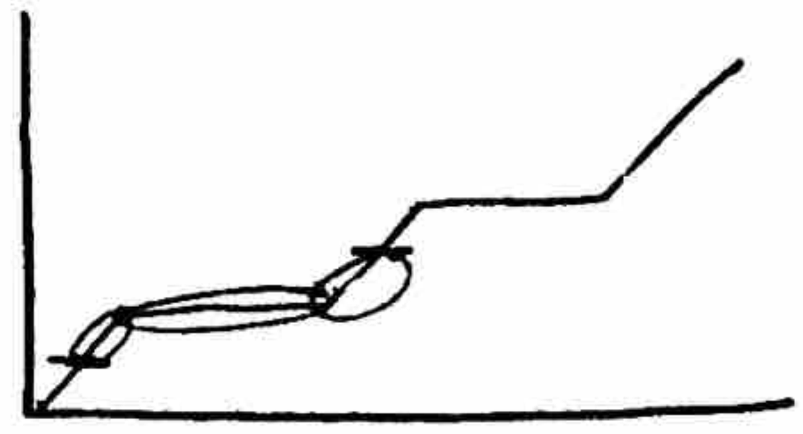
Total Energy in Heating & Cooling Curves

1. Sketch a heating curve of 35 grams of ice at -30°C to water at 57°C, then calculate the total energy required to achieve this.



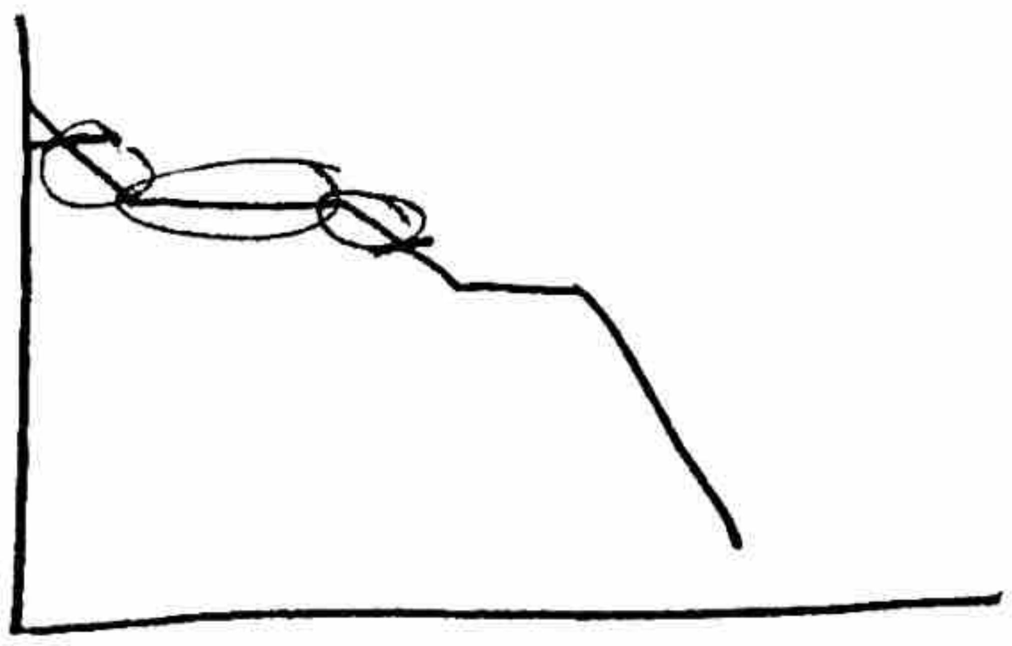
① $q = mc\Delta T = (35g)(2.05\ J/g°C)(30°C) = 2153 \rightarrow 2200$

② $\Delta H = mH_f = (35g)(334\ J/g) = 11690 \rightarrow 12000$

③ $q = mc\Delta T = (35g)(4.18\ J/g°C)(57°C) = 8339 \rightarrow 8300$

22182 J   22500 J

2. Sketch a cooling curve of 452 grams of steam at 231°C condensing to water at 57°C, then determine the amount of energy released by the steam to condense to the water.



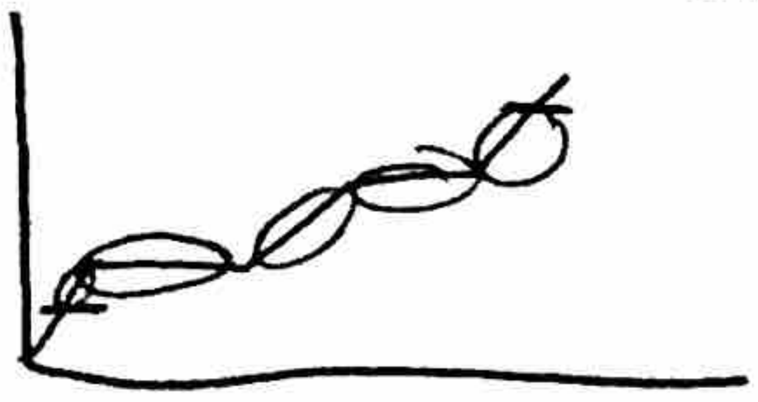
① $q = mc\Delta T = (452g)(2.02\ J/g°C)(100° - 231°C) = -119608 \rightarrow -120000$

② $\Delta H = mH_f = (452g)(-2260\ J/g) = -1021520 \rightarrow -1020000$

③ $q = mc\Delta T = (452g)(4.18\ J/g°C)(57 - 100°C) = -81242 \rightarrow -81200$

1222370 J   -1221200 J

3. Sketch a heating curve of 850.0 grams of ice at -55°C heated to steam at 165°C. Calculate how much energy is required to do this.



① $q = mC\Delta T = (850.0g)(2.05\ J/g°c)(0 - -55) = 96000$

② $\Delta H = mH_f = (850.0g)(334\ J/g) = 283900$

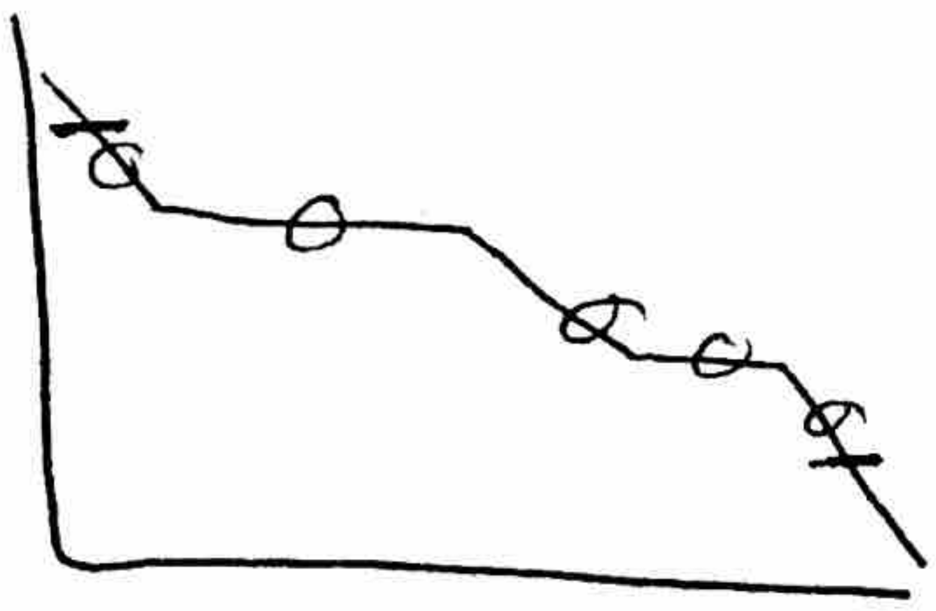
③ $q = mC\Delta T = (850.0g)(4.18\ J/g°c)(100-0) = 355000$

④ $\Delta H = mH_v = (850.0g)(2260\ J/g) = 1921000$

⑤ $q = mC\Delta T = (850.0g)(2.02\ J/g)(165-100) = 110000$

2679500 J

4. Sketch a cooling curve and then calculate the energy lost by 255.0 grams of steam at 123 °C cooled to ice at -21.0 °C.



① $q = mC\Delta T = (255.0g)(2.02)(100 - 123°C) = -12000$

② $\Delta H = mH_v = (255.0g)(-2260\ J/g) = -576300$

③ $q = mC\Delta T = (255.0g)(4.18\ J/g°c)(0-100) = -106600$

④ $\Delta H = mH_f = (255.0g)(-334\ J/g) = -85170$

⑤ $q = mC\Delta T = (255.0g)(2.05\ J/g°c)(-21-0) = -11000$

-791070 J

Multi-step q WS

Name: Key
Date: ____ Period: ____

1. How many joules of heat are required to raise the temperature of 25 g of water from 10°C to 90°C? (8368 J)

$q = mC\Delta T$

$q = (25 g)(4.18 J/g°C)(80°C) = \boxed{8360 J}$ → $\boxed{8000 J}$

2. How many joules of heat are necessary to raise the temp of 42 g of ice from -5°C to liquid water at 25°C? (18854 J)

① $q = (42 g)(2.05)(5) = 430.5 → 400$

② $\Delta H = (42)(334) = 14028 → 14000$

③ $q = (42)(4.18)(25) = 4389 → 4400$

$\boxed{q_{total} = 18847.5 J}$

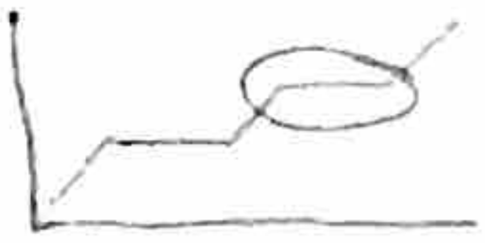
→ $\boxed{18800 J}$

3. How many joules are necessary to melt 50 g of ice? (16700 J)

$\Delta H = (50)(334) = \boxed{16700 J}$

→ $\boxed{20000 J}$

4. How many joules are necessary to vaporize 50 g of water? (113000 J)



$\Delta H = (50)(2260) = \boxed{1.13 \times 10^5 J}$

→ $\boxed{100000 J}$

5. What is the value of q when 60 g of water vapor at 210°C is cooled to 60°C? (-158974 J)

① $q = (60 g)(2.02)(-110) = -13332 → -13000$

② $\Delta H = (60)(-2260) = -135600 → 100000$

③ $q = (60)(4.18)(-40) = 10032 → 10000$

$\boxed{q_{total} = -158964 J}$

→ $\boxed{123000 J}$

6. How many joules of heat are absorbed when 30 g of water at 20°C are heated to 70°C? (6276 J)

$q = (30)(4.18)(50) = \boxed{6270 J}$

→ $\boxed{6000 J}$

7. How much energy is needed to raise the temp of 50 g of ice at -70°C to vapor at 107°C? (158537 J)

① $q = (50)(2.05)(70) = 7175 → 7000$

② $\Delta H = (50)(334) = 16700 → 20000$

③ $q = (50)(4.18)(100) = 20900 → 20000$

④ $\Delta H = (50)(2260) = 113000 → 100000$

⑤ $q = (50)(2.02)(7) = 707 → 700$

$\boxed{q_{total} = 158482 J}$

→ $\boxed{147700 J}$

8. How many joules of heat are necessary to raise the temp of 40 g of vapor from 101°C to 131°C? (2424 J)

$q = (40)(2.02)(30) = \boxed{2424 J}$

→ $\boxed{2000 J}$

9. How much energy is needed to convert 15 g of ice at 0°C to steam at 100°C? (45186 J)

① $\Delta H = (15)(334) = 5010 → 5.0 \times 10^3$

② $q = (15)(4.18)(100) = 6270 → 6000$

③ $\Delta H = (15)(2260) = 33900 → 34000$

$\boxed{q_{total} = 45180 J}$

→ $\boxed{45000 J}$

10. What is q when 35 g of water is vaporized? (79100 J)

$\Delta H = (35)(2260) = \boxed{79100 J}$

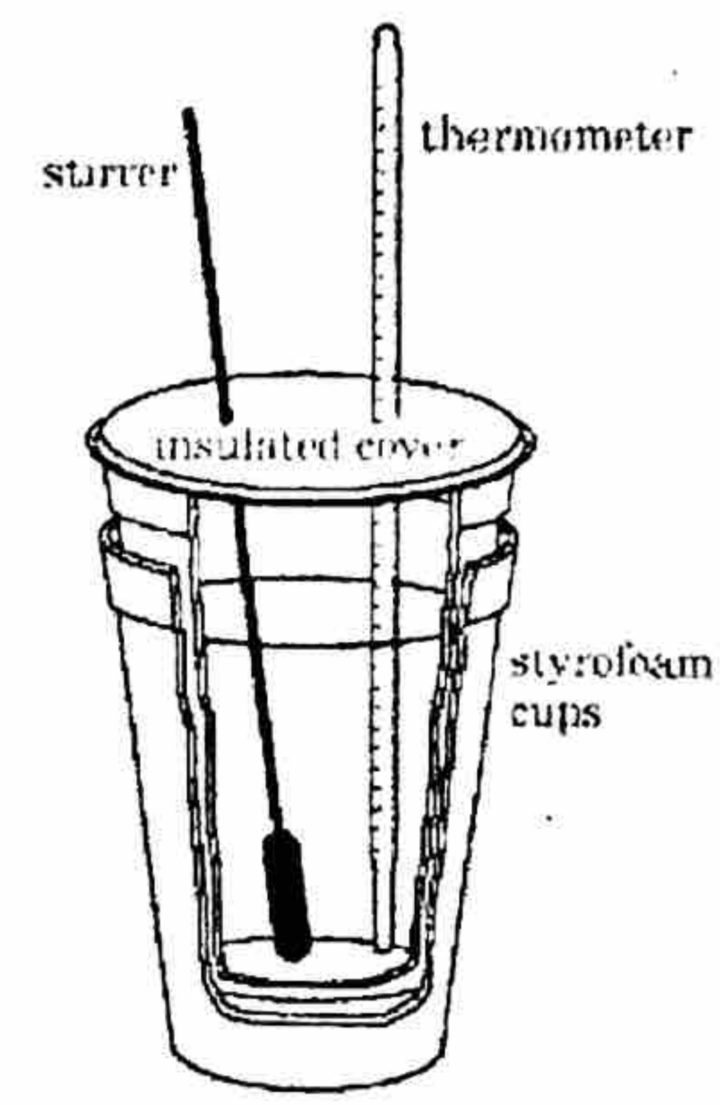
→ $\boxed{79000 J}$

Thermochemistry: Calorimetry

INFORMATION

Calorimetry is the measurement of heat transfer into or out of a system. This measurement can be for physical processes (such as solution formation) or chemical ones (such as chemical reactions). In accordance with the law of conservation of energy, the amount of heat absorbed into the system must be equivalent to the heat released by the surroundings. Conversely, the amount of heat absorbed by the surroundings must be equivalent to the amount of heat released by the system.

An insulated device used to make these measurements is called a **calorimeter**. A constant-pressure calorimeter (one that is open to the atmosphere, if only slightly) is the most common type. In the laboratory, a surprisingly passable and efficient calorimeter can be made using a Styrofoam coffee cup (actually, two nested cups are generally used to ensure good insulation from the surroundings).



Enthalpy (*H*) refers to the heat content of a system at constant pressure, so coffee cup calorimeters are accurate and useful when making measurements of enthalpy. If heat is released or absorbed by the system, the resulting change in enthalpy is represented as ΔH. In constant-pressure reactions (which are the most common type investigated in ~~the laboratory~~), heat and enthalpy are interchangeable terms; that is, $q = \Delta H$.

Enthalpy change for a reaction or process can be calculated simply by utilizing the equation you derived in the previous activity. Recall:

$$q = mc_p\Delta T$$

In this case, q represents the heat absorbed by the surroundings (q_{surr}). To represent the heat absorbed by the system, recall that the sign needs to be inverted:

$$q_{sys} = -q_{surr} = -mc_p\Delta T$$

Recall that the sign is positive (+) for endothermic reactions and negative (-) for exothermic reactions.

Key Questions

1. If you put a piece of hot metal into cold water:
   a. What happens to the temperature of the water? increases
   b. What happens to the temperature of the metal? decreases
   c. What gains heat? water
   d. What loses heat? metal
   e. What has to be true about the heat lost and the heat gained? equal, but opposite
   f. What law tells us that? first law of thermodynamics
   aka
   law of conservation of energy

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Calorimetry

7. A piece of metal with a mass of 4.68 g decreases its temperature from 90.0 C to 30 C when placed into 25g of water at 25 C, what is the specific heat of the metal? (1.9J/g°C)

metal	water
m 4.68g	m 25g
C	C 4.18
Tf 30.0°C	Tf 30.0°C
Ti 90.0°C	Ti 25.0°C

$-(mc\Delta T) = mc\Delta T$

$-(4.68 \cdot x \cdot -60.0) = (25 \cdot 4.18 \cdot 5.0°C)$

$\overline{-1} \qquad \overline{-1}$

$4.68 \cdot x \cdot -60.0 = -522.5$

$x = 1.9\ J/g°C$

8. A piece of metal with a mass of 6.88 g decreases its temperature from 100.0°C to 50.0°C when placed into 15g of water at 40.0°C, what is the specific heat of the metal?

metal	water
m 6.88g	m 15g
C	C 4.18 J/g°C
Tf 50.0°C	Tf 50.0°C
Ti 100.0°C	Ti 40.0°C

$-(6.88g \cdot x \cdot -50.0) = (15g \cdot 4.18 \cdot 10)$

$\overline{-1} \qquad \overline{-1}$

$\frac{6.88 \cdot x \cdot -50}{6.88 \cdot -50} = \frac{-627}{6.88 \cdot -50}$

$x = 1.8\ J/g°C$

9. A 3.55g piece of aluminum, which has a specific heat of 0.897J/g°C, decreases its temperature ~~from~~ to 25°C when placed in 55g of water at 20.0°C. What was the initial temperature of the aluminum? (386°C)

metal	water
m 3.55g	m 55g
C 0.897 J/g°C	C 4.18 J/g°C
Tf 25°C	Tf 25°C
Ti X	Ti 20.0°C

$\frac{-(3.55g \cdot 0.897 \cdot (25-x))}{-3.55 \cdot 0.897} = \frac{(55g \cdot 4.18 \cdot 5°C)}{-3.55 \cdot 0.897}$

$25 - x = -361$

$-25 \qquad -25$

$\frac{-x}{-1} = \frac{-386}{-1}$

$x = 386°C$

10. A 12.6g piece of iron, which has a specific heat of 0.449J/g°C, decreases its temperature to 15°C when placed in 35g of water at 10.0°C. What was the initial temperature of the iron? (144°C)

metal	water
m 12.6g	m 35g
C 0.449 J/g°C	C 4.18 J/g°C
Tf 15°C	Tf 15°C
Ti ?	Ti 10.0°C

$\frac{-(12.6g \cdot .449 \cdot (15-x))}{-(12.6 \cdot .449)} = \frac{(35 \cdot 4.18 \cdot 5)}{-(12.6 \times .449)}$

$15 - x = -129$

$-15 \qquad -15$

$\frac{-x}{-1} = \frac{-144°C}{-1}$

$x = 144°C$