Thermochemistry : Heat and Energy

INFORMATION

Thermochemistry is the study of energy changes and transfers that occur during chemical reactions.

During chemical reactions, energy (in the form of heat) can either be consumed by the reaction (this is called an **endothermic** reaction) or heat can be released by the reaction (this is called an **exothermic** reaction). Energy stored in a compound is really stored in its chemical bonds – this is referred to as **chemical potential energy**. Sometimes, energy is required to break bonds; sometimes, it is required to form them.

Heat, represented by q, is a form of energy. Heat transfers from one object to another naturally as a result of the temperature difference between them. Heat *always* flows from warmer objects to colder ones. Heat will continue to flow until the temperature difference between them is zero.

Exothermic and Endothermic Processes

A **system** is a portion of the universe in which you have an interest. This may seem like somewhat of a daunting statement, but in reality, it simply indicates what you have your attention focused on for the purpose of the problem at hand. For example, a balloon filled with air can be considered a system; a beaker with two aqueous reactants can be a system; a refrigerator is a system; a gigantic chemical plant is a system. Determining what exactly constitutes a system is arbitrary – it is defined based on the needs or interests of the situation.

Surroundings, in contrast, represent everything else in the universe that is not in system. Again, this is broad and seemingly daunting, but usually the surrounding are defined as the area in the immediate vicinity (this may vary by scale – a reaction in a beaker may have "smaller" surroundings than a petrochemical factory, but you get the idea).

Thermochemistry focuses on the study of heat transfer (or **heat flow**) between the system and the surroundings. Exothermic processes result in heat leaving the system and entering the surroundings; endothermic processes result in heat leaving the surroundings and entering the system. Mathematically, heat entering a system has a positive sign ($+q$), and heat entering the surroundings has a negative sign ($-q$).

Recall that the **law of conservation of energy** states that energy can be neither created nor destroyed; therefore, the sum of the total heat in the system and the surroundings must remain the same. This means that as a certain amount of heat leaves the system, the same amount must enter the surroundings (and vice versa).

Heat, Quantitatively

The two common units used for measuring heat are the **calorie** and the **joule**. A calorie is defined as the amount of energy required to increase the temperature of one gram of pure water by one degree Celsius. This applies only to a calorie when written with a small "*c*" in its name. When written with a capital "*C*", as in Calorie, it refers to the more commonly known dietary calorie. Hence:

1000 calories (cal) = 1 kilocalorie (kcal) = 1 Calorie (Cal)

The energy denoted by a calorie, or by a Calorie, is measured by burning a sample under carefully monitored and controlled conditions and measuring the amount of heat released. When a candy bar wrapper indicates that the snack contains 100 Calories, this means that your body can break it down and acquire the equivalent of 100 Calories of energy. If one were to burn the same candy bar and measure the released heat, it would register as 100 Calories (or 100 kcal) of heat. This is, in fact, how Calorie content is determined for food.

The SI unit of energy is the joule. A joule is a smaller unit of energy than a calorie; one joule of heat can only raise the temperature of one gram of water by 0.239 degrees Celsius. Hence:

1 joule (J) = 0.0239 cal **or** 4.184 J = 1 cal

Heat Capacity

Water, iron, and polyethylene plastic all require different amounts of heat to warm them. This property of matter is called **heat capacity**, and is dependent on both the mass of a sample and its chemical composition. A larger sample of water requires more heat to increase its temperature than a smaller sample, for example.

Substance	C_p J/g·C
Asphalt	0.92
Brick	0.84
Concrete	0.88
Glass, silica	0.84
Glass, crown	0.67
Glass, flint	0.503
Glass, pyrex	0.753
Granite	0.79
Gypsum	1.09
Marble, mica	0.88
Sand	0.835
Soil	0.8
Wood	0.42

Early scientists observed that different substances require different quantities of energy to change their temperature. Have you ever noticed that the hood of a car is hot to the touch in the summer sun, but a swimming pool is cool when you jump in? This is because a given mass of a certain substance may require more or less heat than the same mass of another substance to facilitate a change in its temperature. This property of matter (the amount of heat required to raise the temperature of 1 gram of it by 1 degree Celsius) is called **specific heat capacity**, or just **specific heat**. Specific heat must be measured; it cannot be calculated from other observations; it is reported as **J/g·C**, read "joules per gram degree Celsius." Typically, specific heats are looked up on a table of common substances, like the one provided for common substances above. For reference, the specific heat of water is 1 J/g·C.

Key Questions: Answer on your own paper using complete sentences!!

1. An identical exothermic chemical reaction is carried out in two separate beakers. In one beaker, the solvent for the reactants is water; in the other, the solvent is an alcohol. Water and the alcohol have different specific heats. Without measuring, how could you determine which has the higher specific heat after the reaction is completed?

2. Assuming you are given the mass of a substance, the magnitude of desired temperature change, and a table of specific heat capacities, derive an equation that can be used to solve for the total heat required for this process. Recall that ΔT can be used to represent a change in temperature, $T_f - T_i$. *Use the units of the values to make sure the equation works – i.e., the correct units cancel or multiply on one side to equal their configuration on the other.*

Calorimetry – Measurement of Heat Energy

Why?

The amount of energy released or absorbed by a chemical reaction can be measured using a calorimeter. This information is essential to understanding the stability of chemical compounds, predicting equilibrium concentrations in chemical reactions, and identifying conditions for a reaction to occur efficiently and safely. In this activity you will learn how the energy change in a chemical reaction can be measured using a calorimeter.

Success Criteria

- Quantify the relationship between heat absorbed or lost and an observed temperature change.
- Calculate the heat required to raise the temperature of a given mass of water.
- Determine the specific heat capacity of a substance other than water.
- Compare the expected temperature changes for a sample of lead with the expected temperature change for a sample of water of equal mass.

Prerequisites

- Exothermic Change
- Endothermic Change
- Law of Conservation of Energy

Information

- **Heat** is the energy associated with the random motion of particles.
- The unit for **Heat Energy** is the **Joule (J).**
- **Kinetic Energy** is the energy associated with the motion of atoms and molecules.
- **Temperature** is a measure of the **energy** in a sample of material.
- The symbol **ΔT** refers to **"the change in temperature."**
  Example: $\Delta T = 5.00\ ^0C$ means a temperature ***change*** of 5 ^0C.
- **Heat Capacity** is the energy required to raise the temperature of a 1 g sample of a substance 1 ^0C (or 1 Kelvin degree).
- The specific heat capacity for water is 4.18 Joule/gram Kelvin degree

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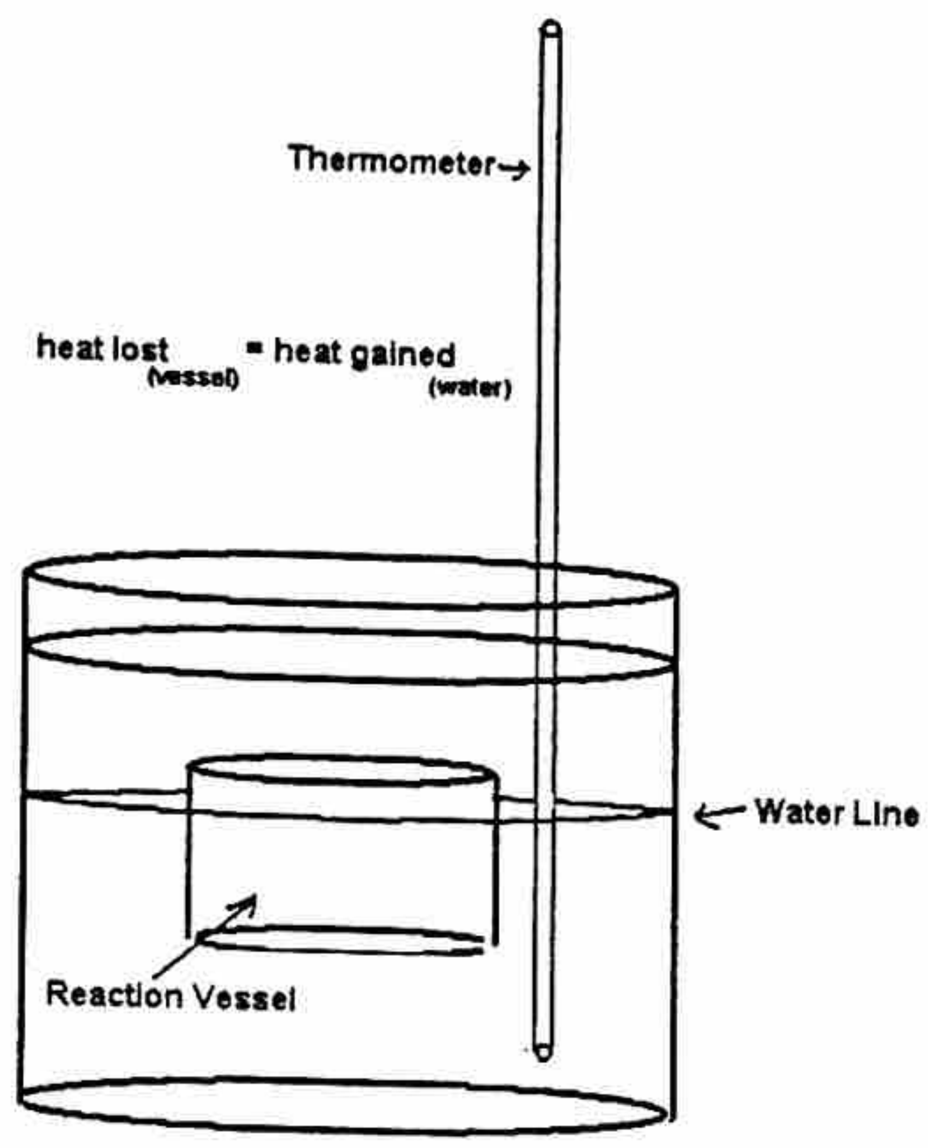
Model

Assume that a calorimeter is a closed system where all the energy released by an exothermic change is absorbed by the water in the calorimeter. If the mass of the water is known, the temperature change of the water can be used to determine the amount of heat energy released.



Calorimeter

Equation for the calculation of heat is: **$q = mC\Delta T$**

q = heat released in Joules
m = mass of water in the calorimeter
C = specific heat capacity (**<u>not ^{0}C which is a temperature</u>**)
ΔT = [final temperature – initial temperature] (absolute value)

Key Questions:

1. What is the numerical value and units of the specific heat capacity of water?

2. What information does the specific heat capacity of water provide?

3. 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.

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Exercises

Answer Exercises 1-5 based upon the passage that follows.

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 ^{0}C. When the reaction was finished the temperature of the water increased to 75.0 ^{0}C.

1. Write the mass of water (m) indicated in the passage.

2. Write the change in temperature (ΔT) indicated in the passage.

3. Write the correct mathematical set-up for the calculation of heat (q)? (Substitute the appropriate values for m, C and ΔT in the equation.)

4. What is the heat quantity released by the chemical change? (Include correct unit)

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

Show the set-up used to solve each of the following exercises. Report your final answer with the correct unit.

6. Determine the heat required to raise the temperature of a 50. g sample of water from 10. ^{0}C to 45 ^{0}C.

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7. Determine the mass of a water sample that is heated from an initial temperature of 25 ^{0}C to a final temperature of 100. ^{0}C following the addition of 1200 J of heat energy.

8. A 100. g sample of pure lead is heated from 10.0°C to 197.5°C by the addition of 3000. J of heat energy. Calculate the specific heat capacity of lead.

Problem

1. The specific heat (C) of water is quite high compared to the specific heat of lead, so the energy released by a sample of heated lead to its surroundings will be only around 4% of the amount of heat released by an equal mass of water under identical conditions.

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

   (b) Calculate the change in temperature that results from the addition of 2500 J of heat energy to a 25g sample of lead.

   (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.

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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)

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

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

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)

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)

Phase Changes

Why?

Most substances go through a phase change when heated or cooled. Molecules of a substance are held together in either the solid, liquid, or gaseous phase by intermolecular forces. It is necessary to discuss what is occurring at the molecular level in order to explain how an ice cube is melted or how water is boiled.

Learning Objectives

- To determine what is occurring on the molecular level during a phase change

Success Criteria

- Students will be able to construct their own heating or cooling curve when given the temperatures at which phase changes occur.

New Concepts

- Heating / cooling curves

Prerequisites

- Temperature (average kinetic energy)
- Phases of matter: solid, liquid, gas
- Molecules
- Endothermic
- Exothermic

Vocabulary

- Fusion (melting)
- Solidification
- Condensation
- Boiling
- Vaporization
- Kinetic energy
- Boiling

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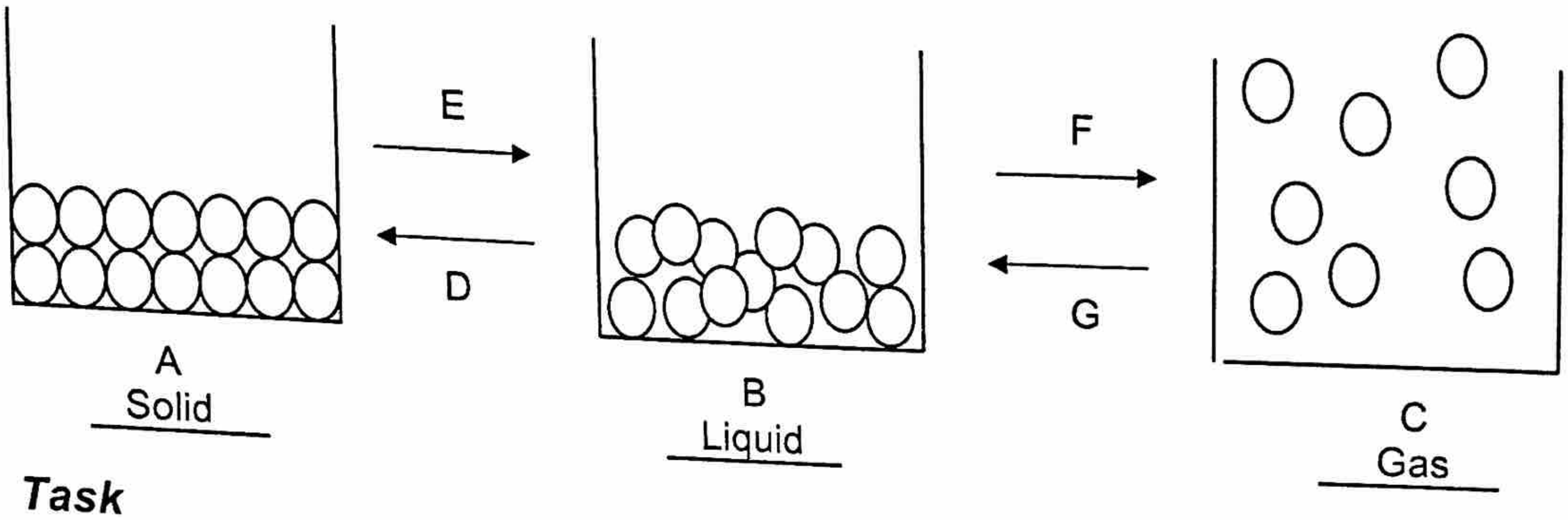
Model 1: Representations of Molecules in Three Phases



Task

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

Key Questions

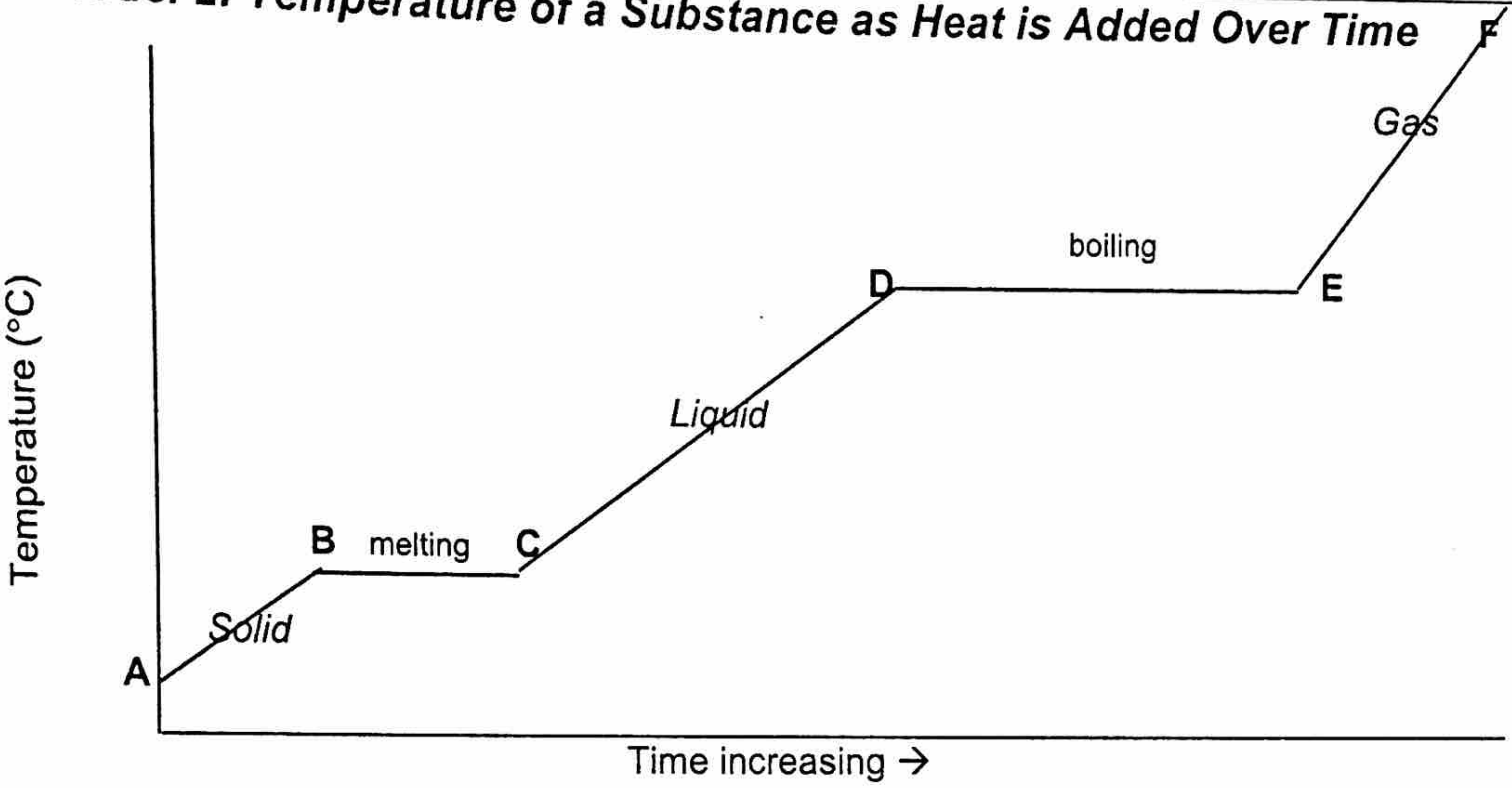
1. Which arrows in Model 1 indicate the addition of energy?

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

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

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

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?

2. During which line segments does temperature increase?

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

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

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

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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.)

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

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

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.

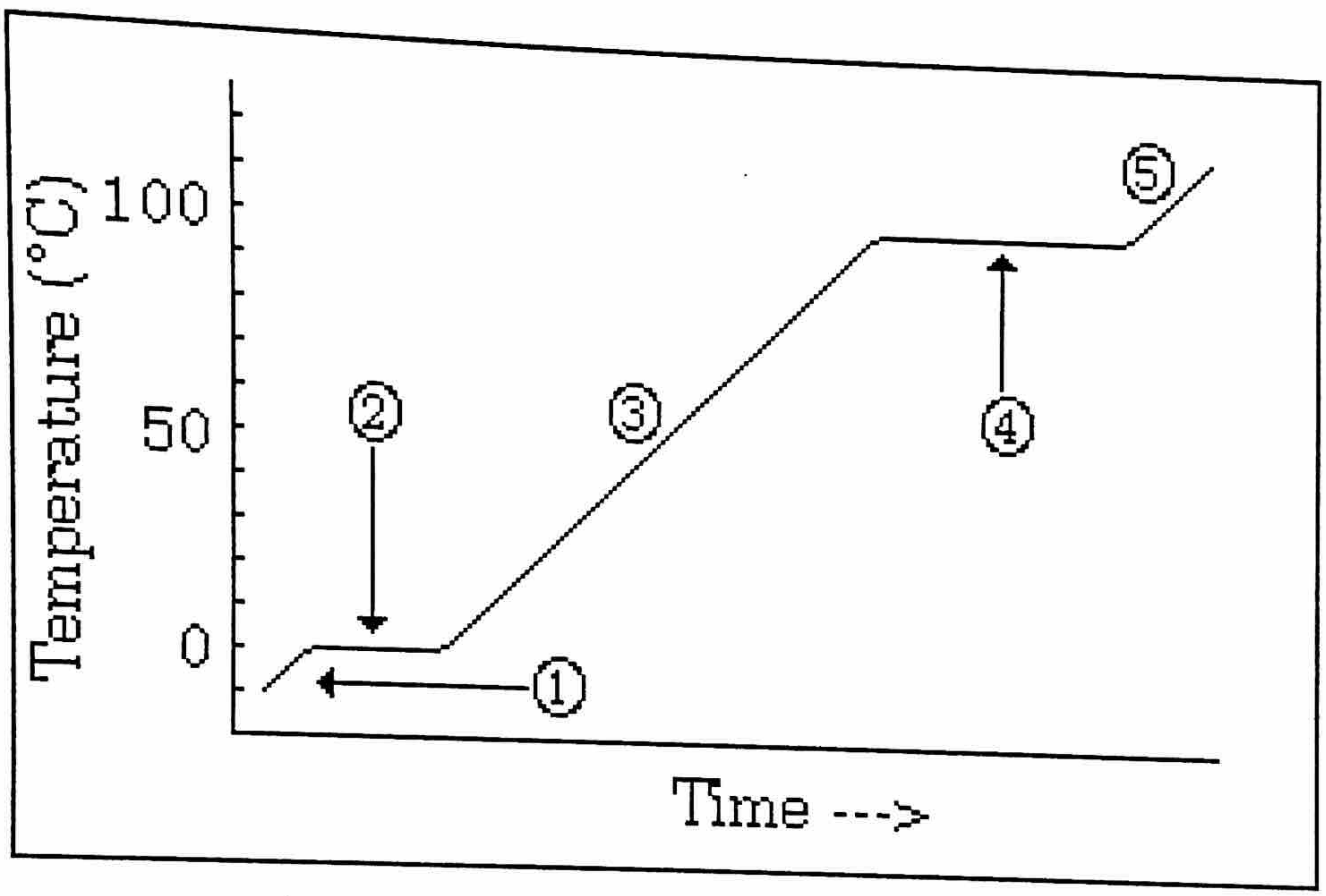
Using this information, draw a cooling curve for the sample.

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

Heating Curve WS #1

The diagram below is a plot of temperature vs. time. It represents the heating of what is initially ice at -10°C at a near constant rate of heat transfer.



1) a)What state(s) of matter are present during segment (1) ____________________

b) What is happening to the particles due to energy being absorbed from the heat source?

____

c) What phase change, if any, is taking place? ____________________

2) a)What state of matter(s) are present during segment (2) ____________________

b) What is happening to the particles due to energy being absorbed from the heat source?

____

c) What phase change, if any, is taking place during segment 2? ____________________

3) a)What state of matter(s) are present during segment (3) ____________________

b) What is happening to the particles due to energy being absorbed from the heat source during segment 3? ________________

c) What phase change, if any, is taking place during segment 3?

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4) a)What state of matter(s) are present during segment 4? ________________

b) What is happening to the particles due to energy being absorbed from the heat source during segment 4? ________________

c) What phase change, if any, is taking place during segment 4? ________________

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5) a)What states of metter are present during segment 5? ________________

b) What is happening to the particles due to the energy being absorbed from the heat source in segment 5? ________________

c) What phase change, if any, is taking place during segment 5? ________________

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6) What is the melting point of this substance?________________

7) At what temperature would this sample boil?____________

8) Circle the spot on the graph where the liquid is melting and boiling.

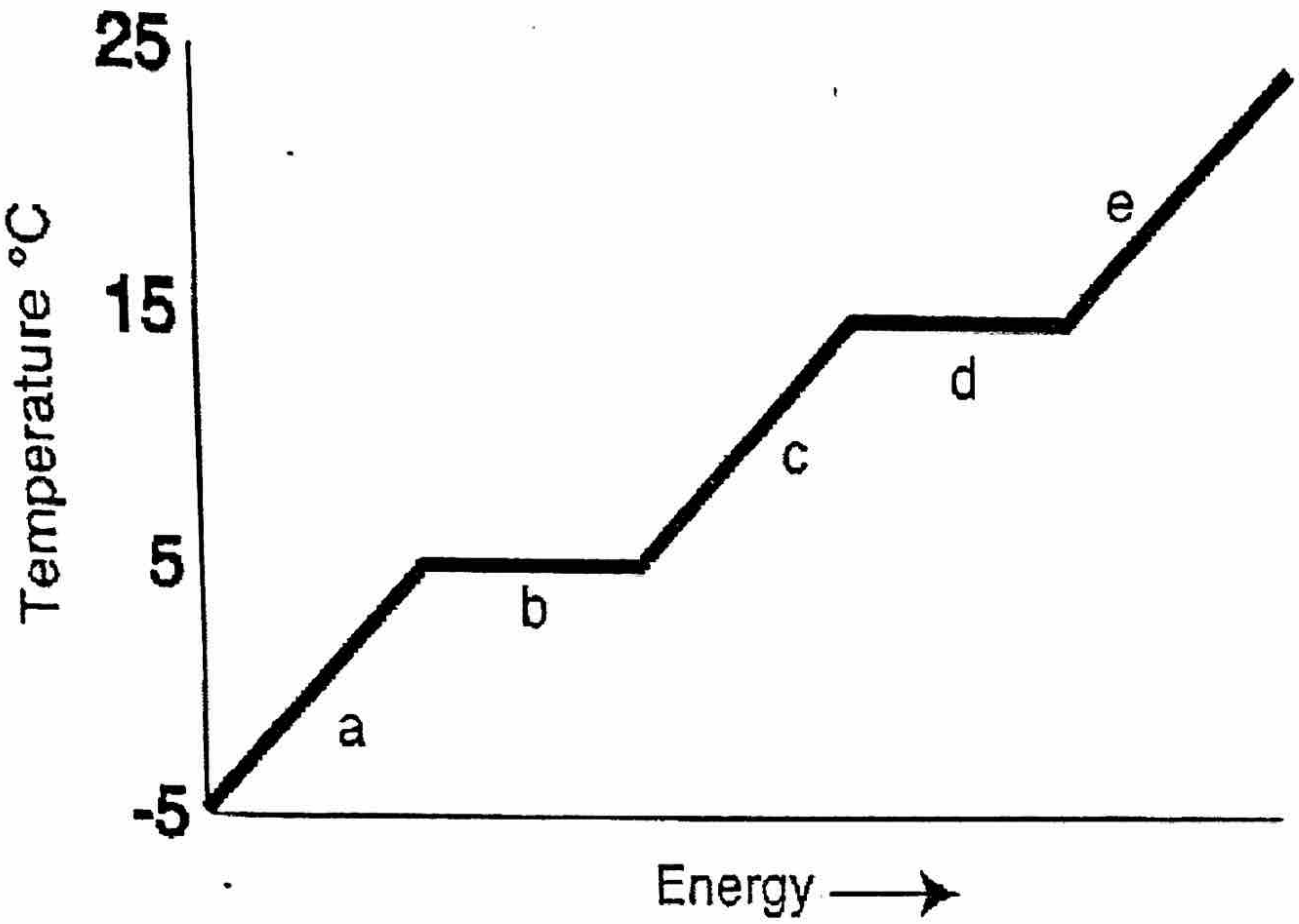
9) Describe the average kinetic energy of the substance as it progresses from #1 to #5.

10) What would the graph look like if it the heat source was much hotter?

Heating Curve Homework

Name________________________

Date______________Per_____



ANSWER THE FOLLOWING USING THE ABOVE HEATING CURVE

1. What is the melting temperature of the above substance? _____
2. What is the freezing temperature of the above substance? _____
3. What is the boiling temperature of the above substance? _____
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. ____

b. ____

c. ____

d. ____

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

Phase Changes

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

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

4. What quantity of ice, at 0°C, will be melted by 6.8 x 10^4 J of energy? (204g)

Total Energy in Heating & Cooling Curves

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

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

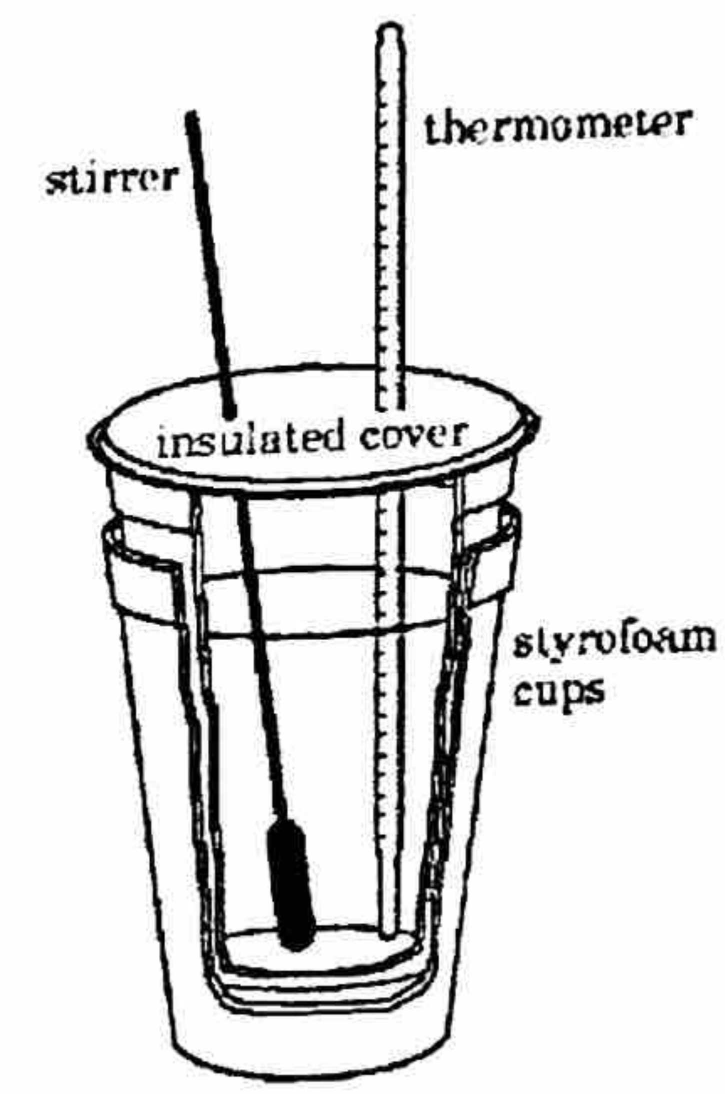
3. Sketch a heating curve of 850.0 grams of ice at -55°C heated to steam at 165°C. Calculate how much energy is required to do this.

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.

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?
   b. What happens to the temperature of the metal?
   c. What gains heat?
   d. What loses heat?
   e. What has to be true about the heat lost and the heat gained?
   f. What law tells us that?

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)

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?

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)

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)