ENERGY

Specific heat and phase changes

WHAT IS ENERGY?

- **Energy**: the ability to do work or produce heat
- Two types of energy
  - **Potential Energy**: energy due to the composition or position of an object (stored energy)
  - **Kinetic Energy**: energy of motion

ENERGY

- Chemical systems contain both kinetic and potential energy.
  - **Potential Energy**: the energy stored in the chemical bonds of a molecule
  - **Kinetic Energy**: the vibration or movement of molecules in a substance
    - This is related to the temperature of an object

IMPORTANT CONCEPT

- An important concept in science is the law that describes energy
- **Law of conservation of energy**: In any chemical reaction or physical process, energy can be converted from one form to another, but it is neither created or destroyed.
  - This concept will become more important when we discuss chemical reactions

TOPICS IN ENERGY

- We will be discussing two aspects of energy in this unit

- Specific Heat
- Changes in States of Matter

WHAT IS HEAT?

- For the next minute, discuss in your groups what you think is the definition of heat.

DEFINITION OF HEAT

- **Heat**: energy that is in the process of flowing from a warmer object to a cooler object
  - Therefore, heat is a relative term
    - 95° is cool when it is summer in Vegas
    - 50° is warm when it is winter in Wisconsin
  - Heat is different from energy and temperature. Again, temperature is a measure of the movement of molecules

UNIT FOR HEAT AND ENERGY

- The unit for both heat and energy are the same. We will be discussing 2 units:
- Calorie (cal)
  - On your food lables, the calorie actually represents 1000 cal or 1 kilocalorie (1 kcal)
- Joule (J)

CALORIE

- **Calorie (cal)**: defined as the amount of energy needed to raise the temperature of one gram of water 1°C
- **Joule (J)**: this is the SI (metric) unit of energy
  - 1 cal = 4.184J

DIMENSIONAL ANALYSIS

- Again, we must use dimensional analysis to convert between units
- REMEMBER: 1 cal = 4.184J
  - This is your conversion
  - Always write down the unit you start with and cancel out the units

EXAMPLE

- You release 250 calories of energy from a chemical sample. How much energy is this in the unit of Joules?

ANSWER

- $\frac{250 \text{ cal} \mid 4.184 \text{ J}}{1 \text{ cal}} = 1046\text{J}$

- REMINDER: Sig figs.
- You start with 2 sig figs. (250)
- You need 2 sig figs. in your answer
  - Therefore: **$1.0\text{x}10^3$ J**

TRY THESE

1. A food item says it contains 2.35×10^3 J of energy. What is this in calories (cal)?
2. You start with 399 cal of energy. What is this in Joules (J)?
3. A yogurt contains 170 calories according to the label. What is this in Joules (J)? (REMINDER: Food label calories are really kilocalories)

WHAT ARE SOME USES OF ENERGY?

- Have you ever noticed that when you boil water, the pot heats up much faster than the water in the pot?
- Why does the pot heat up faster than the water?
- Discuss in your groups for 1 minute why you think this is true.

SPECIFIC HEAT

- The reason is that water needs to absorb more heat than the pot to increase in temperature
- This property of matter is called specific heat
- **SPECIFIC HEAT**: the amount of heat required to raise the temperature of one gram of a substance by 1°C

EXAMPLE

- The specific heat of water is: 4.184 J/(g • °C)
  - We'll discuss the unit in a minute
- The specific heat of concrete is 0.84 J/(g • °C)
- Therefore the specific heat of water is about 5 times bigger than concrete
- **What does this mean?**

ANSWER

- On a hot day, concrete requires less energy to heat
  - Therefore, concrete heats up much faster than water
- This is why you don’t walk on concrete barefoot, but you can still go in the water to cool off
  - Water doesn’t heat up very fast so it stays cooler

HOW HOT DOES SOMETHING GET?

- We defined earlier what specific heat was
- We also showed the units involved in specific heat:
  - **J/(g • °C)**
- We will now discuss how to calculate how hot something gets when exposed to a certain amount of heat or energy

EQUATION FOR HEAT

- To explain temperature change, we have to explain the equation for calculating heat:

$$q = m \times C \times \Delta T$$

DEFINING THE TERMS

- **q**: the heat absorbed or released from a substance in Joules (J) or calories (cal)
- **C**: the specific heat of the substance
- **m**: the mass of the substance in grams (g)
- **ΔT**: the change in temperature in °C (or $T_{final} - T_{initial}$)

EXAMPLE

- Aluminum has a specific heat of 0.897 J/(g • °C). If you add 250 cal of energy to 150g of aluminum, how much does the temperature increase?

STEP 1

- Write down your variables and make sure they are in the proper units:
- $q = 250$ cal
- $c = 0.897$ J/(g • °C)
- $m = 150$g
- $\Delta T = ?$
- NOTE: Heat is in calories and specific heat is in Joules, we need to convert

STEP 1

- $q = 250$ cal
  - $$\frac{250 \text{ cal} \mid 4.184\text{J}}{1 \text{ cal}} = \mathbf{1.0x10^3\ J}$$
- $c =$ **0.897 J/(g • °C)**
- $m =$ **150g**
- $\Delta T = ?$

STEP 2

- Plug your values into the formula:

$$\mathbf{q = m \times C \times \Delta T}$$

$$1.0 \times 10^3 J = 0.897\ J/(g \bullet {}^\circ C) \times 150g \times \Delta T$$

STEP 3

- Perform your calculations as you cancel out units:
- $1.0 \times 10^3 J$ **=** *0.897 J/(g • °C) x 150g* x ΔT
- $\dfrac{1.0 \times 10^3 J}{135\ J/°C} = \dfrac{135\ J/°C \times \Delta T}{135\ J/°C}$
- **7.4 °C = ΔT**

TRY THESE

1. Water has a specific heat of 4.184 J/(g • °C). If you add 125J of energy to 22.2g of water, how much does the temperature increase?
2. Granite has a specific heat of 0.803 J/(g • °C). If you add 125J of energy to 22.2g of granite, how much does the temperature increase?

ANSWER

1. 1.35°C
2. 7.01°C

- Therefore granite heats up about 5 times faster than water

OTHER CALCULATIONS WITH SPECIFIC HEAT

- In addition for solving for temperature change, you can solve for specific heat, amount of heat or the mass of the object.
- Let's explore some examples

EXAMPLE 1

- An object absorbs 1.33×10^4 J of energy. If it has a mass of 15.7g and the temperature increased 25°C, what is the specific heat of the object?
- REMINDER: Make sure you have the correct units in the answer.

ANSWER 1

34 J/g•°C

EXAMPLE 2

- An object has a mass of 35.0g and the temperature increased 15°C. If the object has a specific heat of 0.647 J/g•°C, how much heat was absorbed?

ANSWER 2

340 J

EXAMPLE 3

- A piece of lead has an initial temperature of 25°C. The specific heat of lead is 0.129 J/g•°C. If 55.5g of the lead absorbs 250J of energy, what is the final temperature?

- REMINDER: $\Delta T = (T_{final} - T_{initial})$

ANSWER 3

- $\delta T = 35°C$
- $T_{initial} = 25°C$
- $\Delta T = T_{final} – T_{initial}$
- $35°C = T_{final} - 25°C$

- $\mathbf{T_{final} = 60°C}$

EXAMPLE 4

- An object has a mass of 12.5g and the temperature **DECREASED** 5.7°C. If the object has a specific heat of 0.647 J/g•°C, how much heat was released?
- NOTE: If temperature decreases, you will have a negative "-" ΔT

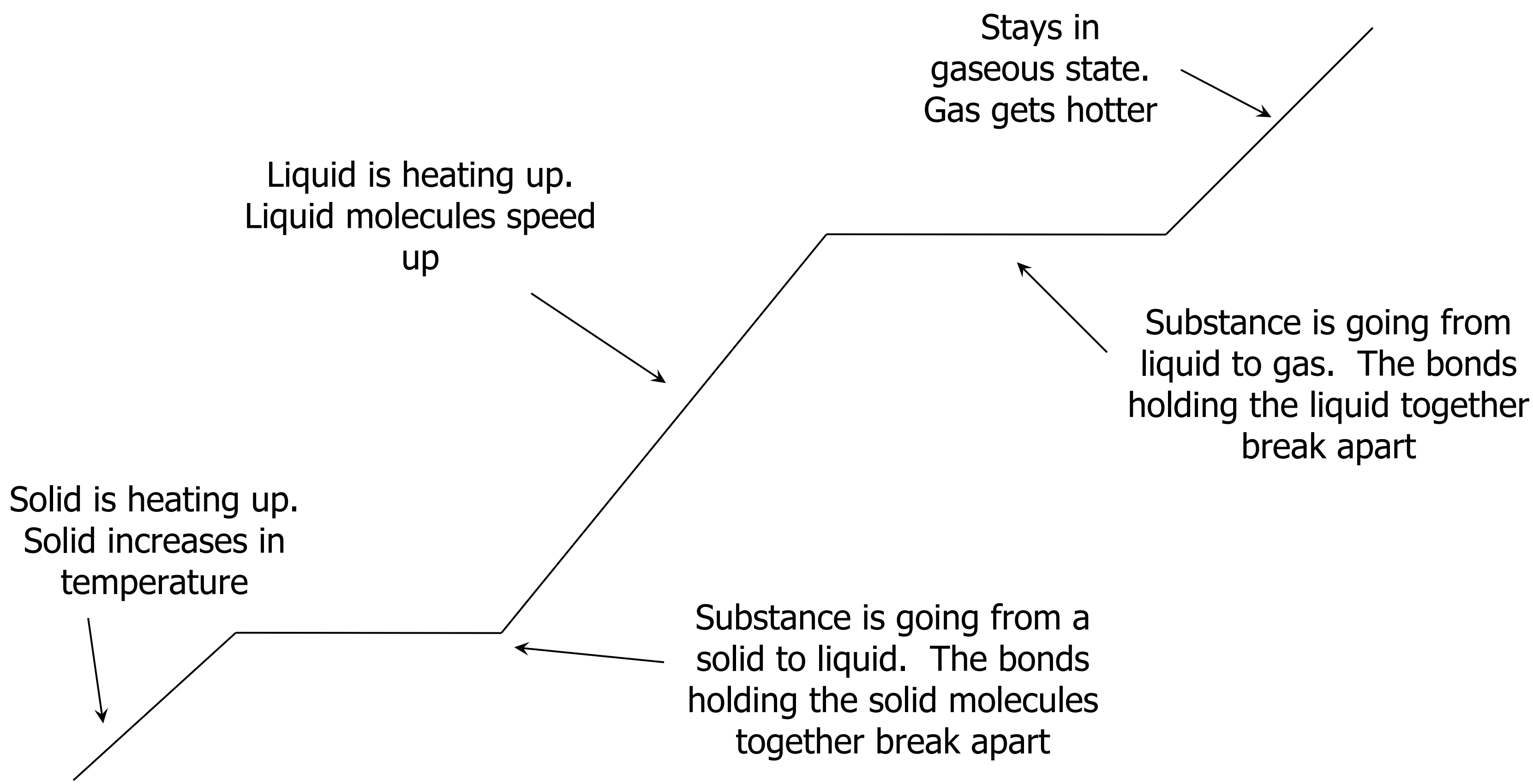
ANSWER 4

-46J

REVIEW

- Earlier in the year, we discussed phase change diagrams
- In your groups, create a phase change diagram:
  - Solid → Liquid;
  - Liquid → Gases

PHASE CHANGE DIAGRAM



SPECIFIC HEAT

- We have already discussed the energy needs for one aspect of the phase change diagram
  - The energy needed to increase the temperature of a substance is dependent on the specific heat
  - SPECIAL NOTE: the specific heat of a substance depending on what state of matter it is in

EXAMPLE

- Specific heats for water:
  - Liquid: 4.184 J/g•°C
  - Solid: 2.03 J/g•°C
  - Gas: 2.01 J/g•°C

CALCULATING HEAT NEEDS

- Knowing the specific heats of substances allows us to determine how much energy is needed to change the temperature of a substance
- What happens to energy needs when there is no change in temperature?
  - EXAMPLE: Phase changes

SOLID → LIQUID

- Before we describe the energy needs for going from solid to liquid, we need to describe how solids and liquids behave at the molecular level

SOLID → LIQUID

- WHAT WE KNOW ABOUT SOLIDS:
  - Solids are close together
  - They are not free moving (bound tightly together)
  - The bonds vibrate, but are fixed
  - Solids tend to be more dense than liquid (this is why solids tend to sink in the liquid of the same matter)
    - EXCEPTION: Water

SOLID → LIQUID

- WHAT WE KNOW ABOUT LIQUIDS:
  - Liquids are still close together
  - They are free moving
  - The bonds that held them together as a solid have been broken apart
  - Liquids are free to move around, but cannot expand further because of intermolecular forces (δ^+ and δ^-)
    - EXAMPLE: Water → Hydrogen bonds

SOLID → LIQUID

- Therefore, to go from a solid to a liquid, you need enough energy to break the bonds holding the solid together
- This is called the heat of fusion
- **<u>Heat of Fusion</u>**: the energy needed to melt one gram of a solid substance

SOLID → LIQUID

- Since there is no temperature change for heat of fusion
  - All the energy is being used to break the bonds holding the solid together
- The formula is:

$$q = m \times H_{fus}$$

EXAMPLE

- You have 250g of ice. The heat of fusion for water is 334 J/g. How much energy is need to melt the ice?
  - $q = ?$
  - $m = 250g$
  - $H_{fus} = 334\ J/g$
- $q = (250g)(334\ J/g) =$ **8.4×10^4 J**

TRY THESE

1. You have 125g of solid acetic acid. The heat of fusion is 195 J/g. What is the heat needed to melt the acetic acid?
2. You have 0.00344g of solid ethanol. The heat of fusion is 107 J/g. What is the heat needed to melt the ethanol?

ANSWERS

1. 2.44×10^4 J
2. 0.368J

LIQUID → GAS

- WHAT WE KNOW ABOUT GASES:
  - Gases are very far apart
  - They are free moving and rapid
  - The bonds that held them together as a liquid have been broken apart
  - Gases do not have any force of attraction holding them together

LIQUID → GAS

- Therefore, to go from a liquid to a gas, you need enough energy to break apart the intermolecular forces holding the liquid together
- This is called the heat of vaporization
- **Heat of Vaporization**: the energy needed to vaporize (or boil) one gram of a liquid substance

LIQUID → GAS

- Since there is no temperature change for heat of vaporization
  - All the energy is being used to break the intermolecular forces holding the liquid together
- The formula is:

$$q = m \times H_{vap}$$

EXAMPLE

- You have 250g of liquid water. The heat of vaporization for water is 2262 J/g. How much energy is need to boil the water?
  - q = ?
  - m = 250g
  - H_{vap} = 2262 J/g
- q = (250g)(2262 J/g) = **<u>5.7×10^5 J</u>**

TRY THESE

1. You have 125g of liquid acetic acid. The heat of vaporization is 390 J/g. What is the heat needed to boil the acetic acid?
2. You have 0.00344g of liquid ethanol. The heat of vaporization is 836 J/g. What is the heat needed to boil the ethanol?

ANSWERS

1. 4.88×10^4 J
2. 28.8 J