### 1. Big Idea:

The enthalpy change of a chemical reaction can be calculated by combining the enthalpy changes of individual steps, thanks to Hess's Law.

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### 2. Essential Questions:

- **How do we calculate the enthalpy change for a reaction using Hess’s Law?**

Answer: Hess’s Law states that the total enthalpy change of a reaction is the same, no matter the number of steps the reaction takes. Using this, the enthalpy change for a reaction can be calculated by summing the enthalpy changes of individual reaction steps if their equations add up to the overall reaction.

- **Why is Hess’s Law useful in thermochemistry?**

Answer: Hess’s Law allows us to calculate the enthalpy change for reactions that cannot be measured directly, providing a way to understand energy transfers even in complex, multi-step reactions.

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### 3. Phenomenon-Based Learning:

### # Unit Phenomenon:

**The Thermodynamics House: Can You Solve the Puzzles and Escape?**

You and your classmates find yourselves trapped in a 2-story high-tech laboratory escape house. To unlock the final door and escape the house, you must solve a series of puzzles presented to you in each room. The puzzles explore how energy flows through chemical reactions and how these reactions behave under different conditions.

### # Chapter Phenomenon:

The second floor of the escape house is devoted to studying the energy involved in chemical reactions. How does energy transfer occur during chemical processes, how much energy is absorbed or released, and how these values can be calculated and manipulated for your escape?

### # Lesson Phenomenon:

**"The Pathway to Escape"**

You have now reached the second to last room of the escape house! Your challenge for this room is to find missing enthalpy values based on provided enthalpy values for each reaction step displayed on interactive screens around the room. Each correct calculation helps piece together clues leading you closer to unlocking your final exit.

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### 4. Vocabulary:

- **Enthalpy:** A measure of the total energy in a system, including both kinetic energy and potential energy stored within chemical bonds.

- **Hess's Law:** A principle stating that the overall enthalpy change for a reaction is equal to the sum of the enthalpy changes for each step in the reaction, as long as the steps combine to form the overall reaction.

- **Standard Enthalpy of Combustion (ΔH<sub>combustion</sub>):** The enthalpy change when one mole of a substance burns completely in oxygen under standard conditions.

- **Standard Enthalpy of Formation (ΔH<sub>formation</sub>):** The enthalpy change when one mole of a compound is formed from its elements in their standard states under standard conditions.

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### 5. SMART Objectives:

- Write thermochemical equations for chemical reactions.

- Calculate the enthalpy change for a reaction and classify it as **endothermic** (absorbing heat) or **exothermic** (releasing heat).

- Use Hess’s Law to determine the enthalpy change for a reaction with multiple steps.

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### 6. Engage (Ignite):

### # Task: Build a Reaction Path Puzzle

Using household items like baking soda, vinegar, and matches, model energy changes in chemical reactions.

**Materials Needed:**

- Baking soda (sodium bicarbonate)

- Vinegar (acetic acid solution)

- A box of matches

- A thermometer

**Procedure:**

1. Measure the temperature of the vinegar in a small cup before adding any baking soda.

2. Add a spoonful of baking soda to the vinegar and stir gently. Record the temperature again.

- **Phenomenon to observe:** The mixture feels cooler. This drop in temperature indicates an **endothermic** reaction.

3. Now light a match and let it burn completely on a plate. Carefully touch the plate (let it cool slightly to avoid burns).

- **Phenomenon to observe:** The plate becomes warmer, signaling an **exothermic** reaction.

**Discussion Questions:**

1. Why does the vinegar-baking soda reaction feel cold?

- Answer: The reaction absorbs energy from its surroundings, making it endothermic.

2. Why does the burned match make the plate warm?

- Answer: The combustion releases energy as heat, making it exothermic.

3. How can these types of energy changes be measured in larger, more complex reactions?

- Answer: By calculating enthalpy changes using data from thermochemical reactions.

**AI Exploration:**

Using an AI tool like ChatGPT, ask it to simulate simple reactions and their enthalpy changes. Example: "What is the enthalpy change when hydrogen reacts with oxygen to form water?" Use this to spark curiosity about larger reactions.

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### 7. Pre-Explore (Direct Instruction):

### # Prior Knowledge:

Chemical reactions involve energy changes due to bond breaking and forming. These changes can either absorb energy (**endothermic**) or release energy (**exothermic**). This concept builds on earlier lessons about energy conservation.

### # Real-World Connection:

Cooking food (endothermic) and burning fuel in cars (exothermic) are everyday processes involving energy changes. Understanding these changes allows for improvements in energy efficiency and reaction control.

### # Background Information:

Hess’s Law helps chemists calculate the energy change for complex reactions that are hard to measure directly. Think of it like finding a shortcut on a road trip: as long as you get from the starting point to the ending point, the total distance (or in this case, total energy change) remains the same.

### # Interactive Elements:

- **Practice Problem:** Consider these two reactions:

1. C (graphite) + O₂ → CO₂ ΔH = −393.5 kJ

2. CO₂ → CO + ½ O₂ ΔH = +283.0 kJ

Use Hess’s Law to calculate the enthalpy change for the overall reaction:

C (graphite) + ½ O₂ → CO.

- **Solution:** Add the enthalpy changes: −393.5 + (+283.0) = −110.5 kJ.

- **Scaffolded Question:** Why can enthalpy changes for steps be added or subtracted?

- Answer: Because enthalpy is a state function, meaning its value depends only on the initial and final states, not the path taken.

This foundation sets the stage for learners to solve multi-step reaction puzzles in the "Explore" phase.

### Chapter 2: The Thermodynamics House – The Pathway to Escape

Welcome to the next stage of the escape room adventure! You’ve reached the second-to-last room on the second floor, and the stakes are high. To move forward, you need to understand how energy flows in chemical reactions and solve enthalpy puzzles. You’ll explore key thermodynamic concepts like **enthalpy**, **Hess’s Law**, and **standard enthalpy** to unlock the door to the final room.

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### Section 1: What is Enthalpy?

Before diving into the puzzles, let’s understand **enthalpy**. Enthalpy (symbol: **H**) measures the total energy of a system, including both internal energy and the energy required to make room for it by displacing the surrounding atmosphere. In simpler terms, it’s the energy stored in bonds of molecules and the energy released or absorbed when those bonds change during a chemical reaction.

Enthalpy change (**ΔH**) is what really matters for chemical reactions. It tells us how much heat energy is absorbed or released during a reaction. This helps us classify reactions into two types:

- **Endothermic reactions**: These absorb energy from their surroundings (ΔH > 0).

- **Exothermic reactions**: These release energy into their surroundings (ΔH < 0).

Imagine you’re baking cookies. The heat absorbed by the dough as it bakes is like an endothermic reaction, while the warmth you feel from the oven is like an exothermic process.

### # Real-World Example: Combustion of Fuel

When you burn gasoline in a car, it releases a large amount of heat energy. This is an **exothermic reaction** because energy is released to do useful work, like moving the car forward.

### # Solved Example:

**Question**: A chemical reaction absorbs 120 kJ of heat energy. What type of reaction is it, and what is the sign of ΔH?

**Solution**: Since the reaction absorbs energy, it’s endothermic, so ΔH = +120 kJ.

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### Section 2: Writing Thermochemical Equations

A **thermochemical equation** is just like a regular chemical equation, but it includes the **enthalpy change (ΔH)**. This shows how energy is involved in the reaction.

For example, when methane (**CH₄**) burns:

\[ \text{CH₄ (g)} + 2 \text{O₂ (g)} → \text{CO₂ (g)} + 2 \text{H₂O (l)} \]

ΔH = -890 kJ (this means 890 kJ of energy is released).

In a thermochemical equation:

- **A negative ΔH** value tells us the reaction is exothermic.

- **A positive ΔH** value tells us the reaction is endothermic.

### # Solved Example:

**Question**: Write the thermochemical equation for the reaction where 50 kJ of heat is absorbed when calcium carbonate (\( \text{CaCO₃} \)) decomposes into calcium oxide (\( \text{CaO} \)) and carbon dioxide (\( \text{CO₂} \)).

**Solution**:

The chemical equation is:

\[ \text{CaCO₃ (s)} → \text{CaO (s)} + \text{CO₂ (g)} \]

Since heat is absorbed, the reaction is endothermic. So, ΔH = +50 kJ.

The thermochemical equation is:

\[ \text{CaCO₃ (s)} → \text{CaO (s)} + \text{CO₂ (g)}, \, ΔH = +50 \, \text{kJ} \]

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### Section 3: Hess’s Law – The Key to Multi-Step Reactions

Now, let’s solve the multi-step reactions in the escape room using **Hess’s Law**. Hess’s Law states:

“The total enthalpy change of a reaction is the same no matter how many steps it takes. It depends only on the starting reactants and final products.”

This is incredibly useful because complicated reactions don’t always happen in one step. Instead, we can add the enthalpy changes of individual steps to find the overall ΔH.

Imagine building a LEGO tower. Whether you build it in 5 minutes or 50 minutes, the tower will still be the same height. Similarly, the total energy change in a reaction stays the same, no matter how many intermediate steps occur.

### # Solved Example:

**Question**:

You’re given two reactions:

1. \[ \text{C (s)} + \text{O₂ (g)} → \text{CO₂ (g)}, \, ΔH = -393.5 \, \text{kJ} \]

2. \[ \text{CO (g)} + \frac{1}{2} \text{O₂ (g)} → \text{CO₂ (g)}, \, ΔH = -283.0 \, \text{kJ} \]

Use Hess’s Law to calculate the enthalpy change for this reaction:

\[ \text{C (s)} + \frac{1}{2} \text{O₂ (g)} → \text{CO (g)} \]

**Solution**:

1. Rearrange reactions to match the target reaction:

- Reverse reaction 2 to get \(\text{CO₂ (g)} → \text{CO (g)} + \frac{1}{2} \text{O₂ (g)}, \, ΔH = +283.0 \, \text{kJ}\).

2. Add the modified reactions:

- \[ \text{C (s)} + \text{O₂ (g)} → \text{CO₂ (g)}, \, ΔH = -393.5 \, \text{kJ} \]

- \[ \text{CO₂ (g)} → \text{CO (g)} + \frac{1}{2} \text{O₂ (g)}, \, ΔH = +283.0 \, \text{kJ} \]

- Total: \[ \text{C (s)} + \frac{1}{2} \text{O₂ (g)} → \text{CO (g)}, \, ΔH = -110.5 \, \text{kJ} \].

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### Section 4: Standard Enthalpy – The Reference Point

To solve tricky puzzles in the escape house, you’ll need to use **standard enthalpy values**. These are measured under specific conditions:

- Temperature: \( 25^\circ \text{C} \) (room temperature).

- Pressure: \( 1 \, \text{atm} \).

The two most important types of standard enthalpy are:

1. **Standard Enthalpy of Formation (\(ΔH\_f^\circ\))**:

This is the energy change when 1 mole of a compound forms from its elements.

Example:

\[ \text{H₂ (g)} + \frac{1}{2} \text{O₂ (g)} → \text{H₂O (l)}, \, ΔH\_f^\circ = -285.8 \, \text{kJ/mol} \].

2. **Standard Enthalpy of Combustion (\(ΔH\_c^\circ\))**:

This is the energy released when 1 mole of a substance burns completely in oxygen.

Example:

\[ \text{CH₄ (g)} + 2 \text{O₂ (g)} → \text{CO₂ (g)} + 2 \text{H₂O (l)}, \, ΔH\_c^\circ = -890 \, \text{kJ/mol} \].

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### Section 5: Unlocking the Puzzle – Practice Problems

You’re almost free from the escape house! Solve these puzzles to escape:

### # Practice Problem 1: Write a Thermochemical Equation

\[ \text{H₂ (g)} + \frac{1}{2} \text{O₂ (g)} → \text{H₂O (l)}, \, ΔH = ? \]

Given that ΔH = -285.8 kJ/mol, write the thermochemical equation.

### # Practice Problem 2: Apply Hess’s Law

Given:

1. \[ \text{N₂ (g)} + \text{O₂ (g)} → 2 \text{NO (g)}, \, ΔH = +180.5 \, \text{kJ} \].

2. \[ 2 \text{NO (g)} + \text{O₂ (g)} → 2 \text{NO₂ (g)}, \, ΔH = -114.1 \, \text{kJ} \].

Find the enthalpy change for:

\[ \text{N₂ (g)} + 2 \text{O₂ (g)} → 2 \text{NO₂ (g)} \].

### # Practice Problem 3: Standard Enthalpy of Combustion

\[ \text{C₃H₈ (g)} + 5 \text{O₂ (g)} → 3 \text{CO₂ (g)} + 4 \text{H₂O (l)} \]

If \( ΔH\_c^\circ \) = -2,220 kJ/mole, how much energy is released when 2 moles of propane burn?

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### Answers to Practice Problems

1. \[ \text{H₂ (g)} + \frac{1}{2} \text{O₂ (g)} → \text{H₂O (l)}, \, ΔH = -285.8 \, \text{kJ} \].

2. Combine the two reactions:

- Total ΔH = \( +180.5 - 114.1 = +66.4 \, \text{kJ} \).

- Final reaction: \[ \text{N₂ (g)} + 2 \text{O₂ (g)} → 2 \text{NO₂ (g)}, \, ΔH = +66.4 \, \text{kJ} \].

3. Energy release = \( 2 \times -2,220 \, \text{kJ} = -4,440 \, \text{kJ} \).

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### Progress Check

1. Define **enthalpy** and explain its importance in chemical reactions.

2. Is a reaction with ΔH = -500 kJ endothermic or exothermic? Why?

3. Use Hess’s Law to calculate ΔH for a two-step reaction if the first step has ΔH = +150 kJ and the second step has ΔH = -200 kJ.

**Answers**:

1. Enthalpy (H) is the total energy of a system, including bond energy and work done to displace surroundings. It shows how energy flows in reactions.

2. Exothermic, because ΔH is negative, meaning energy is released.

3. Total ΔH = \( +150 \, \text{kJ} + (-200 \, \text{kJ}) = -50 \, \text{kJ} \).

By mastering these puzzles, you’re one step closer to escaping the Thermodynamics House! Keep going!

### 9. Elaborate (Power Up)

### # Open-Ended Questions or Mini-Tasks:

1. **What do you think would happen if we lived in a world without chemical reactions?**

- **Answer:** Everyday processes like cooking, breathing, and even growing plants wouldn't occur because these are all driven by chemical reactions. Life as we know it would cease to exist.

2. **How does understanding the structure of atoms help scientists develop new materials?**

- **Answer:** Knowing an atom's structure helps scientists predict and manipulate how atoms bond together. This knowledge leads to the creation of materials with specific properties, like stronger metals or lightweight plastics.

3. **Design an experiment to observe a chemical reaction, like rusting or baking soda reacting with vinegar. What materials would you need, and what would you measure?**

- **Answer:** For a baking soda and vinegar reaction, you would need a measuring cup, vinegar, baking soda, and a balloon. You could measure the gas produced by capturing it in the balloon and noting its expansion.

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### 10. Progress Check (Final Evaluation)

### # Debate Question:

**Should scientists invest more in renewable energy development or focus on improving current fossil fuel technologies?**

**Discussion Points:**

- Renewable energy sources like solar and wind are cleaner and sustainable.

- Improving fossil fuels could reduce emissions and meet energy demands faster in the short term.

**Short Paragraph:**

Renewable energy is gaining attention due to its environmental benefits and sustainability. However, fossil fuels still provide most of the world’s energy, and upgrading current technologies could make them cleaner and last longer. This debate asks us to consider whether we should focus on immediate improvements or long-term solutions for energy needs.

**MCQ:**

- **Which of the following is true about renewable energy?**

- a) It is always cheaper to produce than fossil fuels.

- b) It comes from sustainable sources like the sun and wind. **(Correct Answer)**

- c) It is non-renewable and leads to higher greenhouse gas emissions.

- d) It cannot yet be used to power modern cities.

### # Assessment Questions:

**ACT-Style Multiple-Choice Questions:**

1. **What happens during a chemical reaction?**

- a) Atoms are destroyed.

- b) New atoms are created.

- c) Bonds between atoms are rearranged. **(Correct Answer)**

- d) Energy is always absorbed.

**Explanation:** A chemical reaction involves the rearranging of bonds between atoms to form new substances. Atoms are not created or destroyed in this process.

2. **Which of the following best describes an exothermic reaction?**

- a) Energy is absorbed, and the surroundings feel colder.

- b) Energy is released, and the surroundings feel warmer. **(Correct Answer)**

- c) Energy is absorbed, and the surroundings feel warmer.

- d) It involves no energy change.

**Explanation:** Exothermic reactions release energy into the surroundings, often in the form of heat, making things warmer.

3. **What role does a catalyst play in a chemical reaction?**

- a) It reacts directly with the reactants to form new products.

- b) It slows down the reaction.

- c) It provides energy for the reaction.

- d) It speeds up the reaction without being consumed. **(Correct Answer)**

**Explanation:** A catalyst lowers the energy barrier for a reaction to occur, making it happen faster without being used up in the process.

4. **Which example is NOT a chemical reaction?**

- a) Iron rusting in the presence of water and oxygen.

- b) Baking a cake.

- c) Ice melting. **(Correct Answer)**

- d) Wood burning.

**Explanation:** Ice melting is a physical change because it involves a change of state, not the formation of new substances.

**Long-Answer Questions:**

1. **Describe the difference between physical and chemical changes. Provide one example of each.**

- **Answer:** Physical changes alter a substance's appearance or state of matter but do not create a new substance. For example, ice melting into water. Chemical changes result in the formation of new substances with different properties, such as baking soda reacting with vinegar to produce carbon dioxide.

2. **How does the Law of Conservation of Mass apply to chemical reactions?**

- **Answer:** The Law of Conservation of Mass states that mass is neither created nor destroyed in a chemical reaction. This means the total mass of the reactants equals the total mass of the products. For example, when hydrogen reacts with oxygen to form water, the total mass of the hydrogen and oxygen used is equal to the mass of the water produced.

3. **Explain why balancing chemical equations is important.**

- **Answer:** Balancing equations ensures that the Law of Conservation of Mass is obeyed. It ensures the same number of atoms for each element is present on both sides of the equation so that mass is conserved and the reaction correctly represents what happens in reality.

4. **Design an experiment to test whether a reaction is exothermic or endothermic.**

- **Answer:** Take two beakers, one with vinegar and the other with baking soda. Measure the temperature of both liquids with a thermometer. Mix vinegar and baking soda, and observe the temperature change. If the temperature increases, the reaction is exothermic. If it decreases, the reaction is endothermic.

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### 11. Extend (Beyond the Lesson)

**Tasks, Readings, and Challenges:**

- **Explore:** Research three everyday items that were developed using chemical reactions (e.g., batteries, medicines, or plastics). Write a short paragraph about each.

- **Experiment:** Try growing crystals using a saltwater solution. What observations can you make about how crystals form?

- **Challenge:** Find a news article about a recent discovery in chemistry, like new battery technology or recycling methods, and summarize it for the class.

**Reinforce Understanding Over Time:**

- Revisit the periodic table and find three elements you use in daily life. Write down their symbols and at least one property of each.

- Test yourself by writing balanced chemical equations for simple reactions like combustion or photosynthesis.

**Hints or Connections to Upcoming Topics:**

- Next, we’ll dive deeper into the periodic table! You'll learn how the arrangement of elements reveals patterns and predicts their behavior in chemical reactions. For example, why are sodium and chlorine so eager to bond? Stay tuned to find out!