Sources of Energy Ultimately from the Sun

One of the most common diploma questions involves tracing energy back to its original source. Here, we'll categorize energy sources based on whether they originate from the sun or not.

Solar Energy

The most obvious example of energy coming from the sun is **solar energy**.

> Energy from the sun hits **solar panels**, which convert it to electrical energy. This electrical energy can then be converted into various other forms of energy.

Fossil Fuels

Fossil fuels like gasoline also derive their energy from the sun.

1. Gasoline comes from the ground.
2. Before being in the ground, it was **biomass** (living organic matter) that grew and decomposed.
3. This biomass was covered by rock and sediment.
4. Ultimately, biomass gets its energy from the sun via **photosynthesis**.

Depending on the heat and pressure, biomass can turn into different fossil fuels:

- **Coal:** Solid form when fossil fuels haven't broken down much.
- **Oil:** Liquid form with enough heat and pressure.
- **Natural Gas:** Formed when biomass breaks down even further (e.g., methane).

Biofuels

Instead of waiting for biomass to decompose over millions of years, we can use **biofuels**.

> Fuels derived from recently living organic matter (biomass).

Growing corn or sugar and turning it into fuel is an example. The energy in biofuels comes from the sun via photosynthesis.

Wind Energy

Wind energy also originates from the sun.

1. Windmills harness the kinetic energy of the wind.
2. The wind blows due to the sun heating the earth unequally.
3. This unequal heating causes air masses to rise at different rates, creating wind.

Hydroelectric Energy

Hydroelectric energy is another form of solar energy.

1. Incoming solar radiation drives the water cycle.
2. Evaporation leads to moisture at high points.
3. The potential energy of this water is transferred into electrical energy.

Sources of Energy Not from the Sun

Not all energy on Earth comes from the sun. Here are a few examples:

Nuclear Energy

Nuclear energy comes from the fission or fusion of atoms, which releases energy.

> **Nuclear Fission:** The splitting of a heavy, unstable nucleus into two lighter nuclei.
>
> **Nuclear Fusion:** The process where two light nuclei combine to form a single heavier nucleus.

Tidal Energy

Tidal energy harnesses the kinetic energy of the ocean's tides (high and low points). The tides are driven by the moon, not the sun.

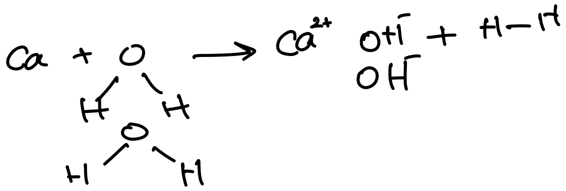
Geothermal Energy

Geothermal energy uses the heat of the Earth's core, which doesn't come from the sun.

Defining Calorimetry

We will discuss calorimetry and the energy changes in chemical reactions. For example, consider the balanced chemical equation:

$$Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2$$



In this reaction, we're interested in the energy change.

Potential Energy of Bonds

The energy change in a chemical reaction is due to the change in potential energy of the bonds.

- At the beginning of the reaction, we have calcium metal with metallic bonds and water molecules with oxygen-hydrogen (O-H) bonds.
- As the reaction proceeds, the calcium loses electrons to become Ca^{2+}, and the water molecules are broken apart into hydroxide ions (OH^-). Energy is required to break the O-H bonds. The hydrogen atoms then form a covalent bond to create H_2 gas.

The change in energy (ΔE) is the difference in potential energy between the bonds at the beginning and the end of the reaction. If there is less energy at the end, the extra energy is released into the environment. If there is more energy at the end, the energy comes from the chemical environment.

more → less : exo
less → more : endo

Measuring Energy Change with Calorimetry

Calorimetry is the measure of energy change in an isolated system. A calorimeter is the device used to make these measurements.

> Calorimetry: The process of measuring the amount of heat released or absorbed during a chemical reaction or physical change.

A typical calorimeter consists of:

- One or two styrofoam cups
- A lid
- A thermometer

The styrofoam cups provide an isolated system, preventing energy from escaping or entering.

Energy Transfer in the Calorimeter

If we know the energy change inside the calorimeter, we can determine the energy change in the chemical reaction. The chemical reaction is the system, and the potential energy change is related to the system. The surroundings, often water, receive the energy.

> System: The chemicals involved in the chemical reaction.
>
> Surroundings: The environment around the chemical reaction, often water, where energy is transferred. → often water

We measure the temperature change in the surroundings to determine the energy change.

Calculating Energy Change

The energy change (q) can be calculated using the formula:

$$q = mc\Delta T$$

Where:

- m = Mass of the warming substance (often water) g
- c = Specific heat capacity of the substance J/g°C
- ΔT = Change in temperature °C

Specific Heat Capacities of Common Substances

Substance	Specific Heat Capacity (J/g°C)
Water	4.19
Other	See Chem 30 Data Booklet

Heat Transfer and Enthalpy Change

Enthalpy (H) is the total kinetic plus potential energy within a chemical system, often simplified to just potential energy in Chem 30. We primarily care about the **change in enthalpy** (ΔH) from reactants to products.

> Enthalpy: The total heat content of a system at constant pressure. It is the sum of the internal energy of the system plus the product of its pressure and volume.

Calculating Enthalpy Change

1. Calculate the **energy transferred to the surroundings** using $q = mc\Delta T$.
2. Determine the **sign of ΔH**:
    $A+B \longrightarrow C + Energy \Rightarrow \Delta H = -value$
    - **Negative** when energy is released (exothermic reaction) : Energy is considered a product ⟶ right-side
    - **Positive** when energy is absorbed (endothermic reaction): Energy is considered a reactant ⟶ left-side
    $A+B+Energy \longrightarrow C \Rightarrow \Delta H = +value$

Assumptions in Calorimetry

To solve calorimetry problems, we make a few key assumptions:

1. **Isolated System**: Assume the calorimeter doesn't gain or lose energy from the surroundings.
2. **Constant Specific Heat Capacity of Water**: Assume the specific heat capacity of water is constant at 4.19 J/g°C.
3. **Specific Heat Capacity of Dilute Aqueous Solutions**: Assume the specific heat capacity of dilute aqueous solutions is the same as water. very important
4. **Density of Dilute Aqueous Solutions**: Treat the density of dilute aqueous solutions as the same as water (1.0 g/mL).

Dilute Aqueous Solutions

For dilute aqueous solutions, a one-to-one ratio is used to determine mass. For example, 100.1 grams of HCl aqueous solution is assumed to be 100.1 mL.

- The more concentrated the aqueous solution, the more this assumption becomes problematic.
- In lab experiments involving mixing a dilute acid and a dilute base, these assumptions are used.

Solving for Enthalpy Change

To solve for enthalpy change ($DeltaH$), use the formula:

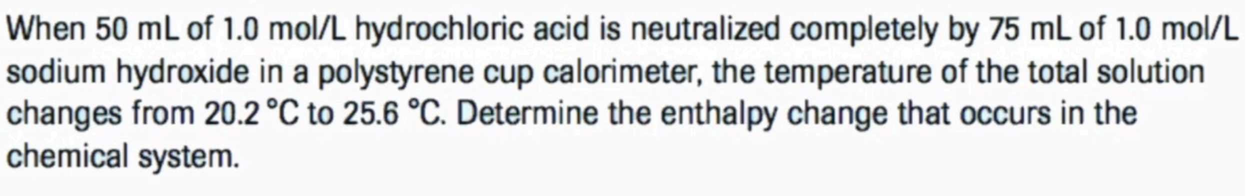
$$\Delta H = mC\Delta T$$

Where:

- m is the mass: g
- c is the specific heat capacity: $\frac{J}{g°C}$
- $DeltaT$ is the change in temperature: °C or k

This calculation focuses on the surroundings, which is often water in these scenarios.

Example Problem



When 50 mL of 1.0 mol/L hydrochloric acid is neutralized completely by 75 mL of 1.0 mol/L sodium hydroxide in a polystyrene cup calorimeter, the temperature of the total solution changes from 20.2 °C to 25.6 °C. Determine the enthalpy change that occurs in the chemical system.

Solution: $Hcl + NaOH \rightarrow NaCl + H_2O$

0.05L = 50mL $\frac{1 mol}{L}$ 75mL = 0.075L $\frac{1 mol}{L}$

volume = 50 + 75 = 125 mL = 125g

$Q = mC\Delta T = 125 \times 4.19 \times 5.4 = 2828J = \boxed{2.8kJ}$

System is getting warmer → exothermic → negative enthalpy

$\Delta H = -2.8kJ$

1. **Identify the knowns:**
    - Volume of HCl: 50 mL
    - Volume of $NaOH$: 75 mL
    - The solvent is water (if not specified, assume it's water).

2. **Determine the mass of water:**
    - The total volume of water is the sum of the volumes of HCl and $NaOH$:

    $$75mL + 50mL = 125mL$$

    - Convert the volume of water to mass, using the density of water ($1g/mL$):

    $$125mL = 125g$$

3. **Find the specific heat capacity of water:**
    - $c = 4.19J/gcdot^{circ}C$ (from data booklet)

4. **Calculate the temperature change ($DeltaT$):**
    - $DeltaT = T_{final} - T_{initial}$
    - Using example values of $T_{final} = 25.6^{circ}C$ and $T_{initial} = 20.2^{circ}C$:

    $$DeltaT = 25.6^{circ}C - 20.2^{circ}C = 5.4^{circ}C$$

5. **Solve for the heat absorbed by the water:**
    - Using the formula:

    $$q = mcDeltaT$$

    - Plugging in the values:

    $$q = (125g)(4.19J/gcdot^{circ}C)(5.4^{circ}C) = 2800J$$

    - Convert joules to kilojoules:

    $$2800J = 2.8kJ$$

6. **Determine the sign of $DeltaH$:**

- If the final temperature is warmer, the reaction is exothermic, and $DeltaH$ is negative.
- If the final temperature is colder, the reaction is endothermic, and $DeltaH$ is positive.
- In this example, if the final temperature is warmer, then:

$$DeltaH = -2.8kJ$$

- The negative sign indicates that energy is released from the chemicals to the surroundings (water).

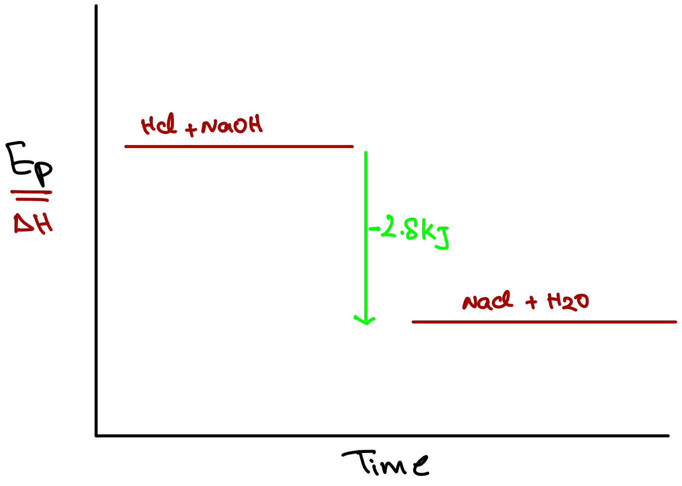
Chemical Potential Energy Diagram

A chemical potential energy diagram illustrates the change in potential energy (a proxy for enthalpy change) during a chemical reaction.

- **Potential Energy:** Energy stored within a system that has the potential to do work.

  > In the context of chemistry, potential energy refers to the energy stored in the bonds between atoms and molecules.

- The diagram plots potential energy versus time.
- In the example of neutralizing HCl with $NaOH$, the products (water and sodium chloride) have lower potential energy than the reactants.
- The difference in potential energy between reactants and products is $2.8kJ$, which corresponds to the energy released into the water.
- The negative sign indicates a loss of potential energy, meaning the chemical bonds in the products have less potential energy than in the reactants. This energy is released as heat, warming the surroundings.



Molar Enthalpy and Calorimetry

Definition of Molar Enthalpy

> Molar enthalpy is the enthalpy change per mole of chemical that reacts.

Key Principle

The change in enthalpy (ΔH) increases as more limiting reagent is added. This is analogous to adding more wood to a fire to get more heat.

Molar Enthalpy Formula

The ratio of the change in enthalpy to moles is a constant, leading to the formula:

$$\frac{\Delta H}{n} = \Delta H_m$$

Where:

- ΔH is the change in enthalpy (typically in kilojoules, kJ)
- n is the number of moles
- ΔH_m is the molar enthalpy (in kJ/mol)

For a given reaction type, every chemical has a constant molar enthalpy. For example, methane releases a fixed amount of kilojoules for every mole that reacts.

Calculating Molar Enthalpy

Steps to Calculate Molar Enthalpy

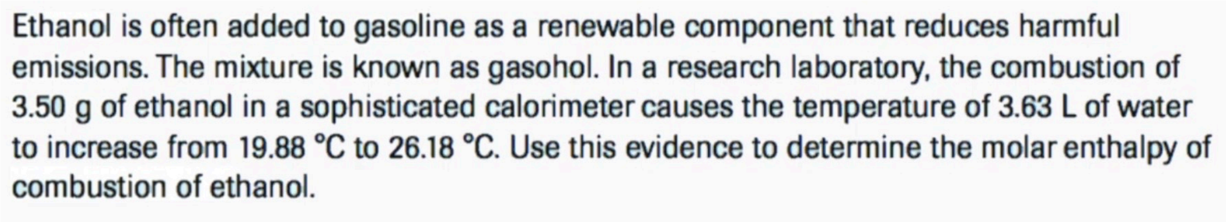
1. Determine the amount of heat (q) or thermal energy transferred to the surroundings (often water) using the formula: $q = mc\Delta T$
   - Where:
     - m is the mass
     - c is the specific heat capacity
     - ΔT is the change in temperature
2. Determine the sign of ΔH:
   - Positive (+) for endothermic reactions (surroundings cool down).
   - Negative (-) for exothermic reactions (surroundings warm up).
3. Divide the enthalpy change (ΔH) by the moles (n) of the limiting reagent: $\Delta H_m = \frac{\Delta H}{n}$

Calculating Moles

- For a pure substance: $n = \frac{mass}{molarmass}$
- For a solution: $n = concentration \cdot volume$

Example Calculation

Problem



Ethanol is often added to gasoline as a renewable component that reduces harmful emissions. The mixture is known as gasohol. In a research laboratory, the combustion of 3.50 g of ethanol in a sophisticated calorimeter causes the temperature of 3.63 L of water to increase from 19.88 °C to 26.18 °C. Use this evidence to determine the molar enthalpy of combustion of ethanol.

$C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$ molarE = ?

3.50g → mol

3.63L ×1000

26.18 − 19.88 = 6.3

Q = 3.63g × 4.19 × 6.3 = 95.8211 kJ

$$3.50\,g \times \frac{1\ mol}{\cancel{24.02+6.06+16}\ 46.08} = 0.076$$

0.076 mol 95.82 kJ

1 mol 1260 kJ/mol

Combustion : Exo:

−Enthalpy

$\Delta H/mol = -1260\ kJ/mol$

1. **Calculate the heat transferred (q) to the water:**
    - Convert volume of water to mass:
        1. 63 L = 3630 mL = 3630 g (since 1 mL of water = 1 g)
    - Use $q = mc\Delta T$: $q = (3630, \mathrm{g}) \cdot (4.19, \frac{\mathrm{J}}{\mathrm{g}\cdot{}^\circ\mathrm{C}}) \cdot (26.18, {}^\circ\mathrm{C} - 19.88, {}^\circ\mathrm{C})$ $q = (3630, \mathrm{g}) \cdot (4.19, \frac{\mathrm{J}}{\mathrm{g}\cdot{}^\circ\mathrm{C}}) \cdot (6.30, {}^\circ\mathrm{C}) = 95821, \mathrm{J} = 95.821, \mathrm{kJ}$
2. **Determine the sign of ΔH:**
    - Since the final temperature is warmer, the reaction is exothermic, so ΔH is negative.
    - $\Delta H = -95.821, \mathrm{kJ}$
3. **Calculate the moles of ethanol:**
    - Molar mass of ethanol (C_2H_5OH) = 46.08 g/mol
    - $n = \frac{mass}{molarmass} = \frac{3.5,\mathrm{g}}{46.08,\frac{\mathrm{g}}{\mathrm{mol}}} = 0.0760, \mathrm{mol}$
4. **Calculate molar enthalpy (ΔH_m):** $\Delta H_m = \frac{\Delta H}{n} = \frac{-95.821,\mathrm{kJ}}{0.0760,\mathrm{mol}} = -1260, \frac{\mathrm{kJ}}{\mathrm{mol}}$
5. **Round to three significant figures:** $\Delta H_m \approx -1260, \frac{\mathrm{kJ}}{\mathrm{mol}}$

Answer

The molar enthalpy for the combustion of ethanol is approximately -1260 kJ/mol. This value is a constant for the combustion of ethanol and can be compared to data tables.

Calorimetry Wrap-Up: System vs. Surroundings

It's important to understand the distinction between the system and the surroundings in calorimetry:

- System: Refers to the reactants or chemicals involved in the reaction. This is where you need to identify the limiting reagent to calculate the number of moles. The system is defined by the chemical equation. We cannot directly measure the change in chemical bonds (ΔH) within the system.

- Surroundings: Typically the solution, often water, where $Q = mc\Delta T$ is measured. Sometimes, you may need to calculate Q multiple times if heat is absorbed by multiple components (e.g., water and the metal of the calorimeter).
  - For styrofoam calorimeters, you primarily focus on Q for the water (Q_{water})
  - For metal can calorimeters, you may need to calculate Q for both the water and the metal can ($Q_{total} = Q_{water} + Q_{metal}$).

Enthalpy Subscripts

When discussing enthalpy, we use specific subscripts to provide more information:

- r: Generic reaction that doesn't fit combustion, formation, or decomposition. Analogous to "c" in K_c (equilibrium constant).
- c: Combustion reaction. Indicates molar enthalpy of combustion.
- f: Formation reaction. Indicates molar enthalpy of formation (often found in data books).
- d: Decomposition reaction. Indicates molar enthalpy of decomposition.

When writing these values, you need to specify the change in enthalpy and the molar values for the type of reaction.

Formation Reactions

> Formation is the creation of a compound from its elements in their standard states.

For example, the formation of water (H_2O) can be written as:

$$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$$

Elemental Forms: Quick Recap

Element	Elemental Form	Notes
Most Metals	Single atoms (e.g., Na, Fe)	Except for mercury (Hg)
Mercury	Hg	Exists as a liquid at room temperature
Diatomic Elements	$H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2$	"Seven up" mnemonic: Start at atomic number 7 (Nitrogen) and form a "7" on the periodic table
Phosphorus	P_4	
Sulfur	S_8	Most common form
Carbon	C	Exists in various allotropes (e.g., graphite, diamond)

Decomposition Reactions

> Decomposition is the breakdown of a compound into its elements.

For example, the decomposition of sodium chloride ($NaCl$) can be written as:

$$2NaCl(s) \rightarrow 2Na(s) + Cl_2(g)$$

Communicating Enthalpy: Molar Enthalpy

> Molar enthalpy is the amount of energy released or absorbed per mole of a substance in a reaction.

It is often quoted at Standard Ambient Temperature and Pressure (SATP):

- Temperature: 25°C
- Pressure: Approximately 1 atmosphere
- Substance: Must specify the limiting reagent and reaction type.

Symbolic Representation

The general form for representing molar enthalpy includes:

ΔH_m (change in enthalpy) with:

- Subscript to denote the reaction type (r, c, f, d)
- Chemical formula of the substance involved.
- "°" symbol to indicate standard state (SATP).

For example, the combustion of methane (CH_4) at SATP can be represented as:

$\Delta H_{m,c,CH_4}$

Enthalpy Change and Balanced Equations

Enthalpy change must always be associated with a balanced chemical equation. The value is specific to the reaction as written.

$$CO(g) + 2H_2(g) \rightarrow CH_3OH(l) \quad \Delta H = -128, kJ$$

This means that 128 kJ of energy is released when one mole of carbon monoxide reacts with two moles of hydrogen to produce one mole of methanol.

If the equation is scaled (e.g., doubled):

$$2CO(g) + 4H_2(g) \rightarrow 2CH_3OH(l) \quad \Delta H = -256, kJ$$

Doubling the reaction doubles the amount of energy released. It's like baking: double the ingredients, double the cookies.

Determining Molar Enthalpy from Enthalpy Change

You can calculate the molar enthalpy of a specific chemical using the relationship:

$$\mathit{\Delta} H_m = \frac{\mathit{\Delta} H}{n}$$

Where:

- ΔH_m is the molar enthalpy
- ΔH is the change in enthalpy for the reaction
- n is the number of moles of the substance

For example, given:

$$CO(g) + 2H_2(g) \rightarrow CH_3OH(l) \quad \Delta H = -128, kJ$$

The molar enthalpy of carbon monoxide (CO) is:

$$\Delta H_{m,CO} = \frac{-128, kJ}{1, mol} = -128, kJ/mol$$

Stoichiometric coefficients are exact numbers, not subject to significant figures.

Molar Enthalpy Calculations

Molar enthalpy for hydrogen gas (H_2) is calculated as follows:

$-128, \text{kJ}$ per 2 moles of H_2.

Molar enthalpy: $\frac{-128, \text{kJ}}{2, \text{mol}} = -64.0, \text{kJ/mol}$

Writing Balanced Chemical Equations with Enthalpy

To write a balanced equation with enthalpy, consider the reaction of sulfur dioxide (SO_2) with oxygen (O_2) to form sulfur trioxide (SO_3).

$$SO_2 + O_2 \rightarrow SO_3$$

Balancing this equation gives:

$$2SO_2 + O_2 \rightarrow 2SO_3$$

Given the molar enthalpy of SO_3 is $-98.9, \text{kJ/mol}$, the enthalpy change (ΔH) for the reaction can be calculated using the formula:

$$\Delta H = \text{molar enthalpy} \times \text{moles}$$

For this reaction:

$$\Delta H = -98.9, \text{kJ/mol} \times 2, \text{mol} = -198, \text{kJ}$$

Therefore, the balanced equation with enthalpy is:

$$2SO_2 + O_2 \rightarrow 2SO_3 \quad \Delta H = -198, \text{kJ}$$

Energy Terms in Balanced Chemical Equations

Instead of writing the enthalpy change next to the balanced chemical equation, it can be included directly in the equation as a reactant or product.

For an **exothermic** reaction, energy is released, so it is written as a product. For example, the reaction of sulfur dioxide to sulfur trioxide releases energy:

$$2SO_2 + O_2 \rightarrow 2SO_3 + 198, \text{kJ}$$

In this case, the negative sign indicates energy is released, but in the equation, energy is added as a positive term on the product side.

For an **endothermic** reaction, energy is absorbed, so it is written as a reactant. Consider the formation of ethene (C_2H_4) and hydrogen (H_2) from ethane (C_2H_6):

$$C_2H_6 + 136, \text{kJ} \rightarrow C_2H_4 + H_2$$

Here, energy is required to break the bonds in ethane and form ethene and hydrogen.

Chemical Potential Energy Diagrams

Chemical potential energy diagrams illustrate the energy changes during a chemical reaction. They plot the potential energy (or enthalpy) of the system against the reaction coordinate.

Key components:

- **Reaction Coordinate:** Represents the progress of the reaction (x-axis). No units.
- **Potential Energy/Enthalpy:** Energy of the system (y-axis), typically in kilojoules (kJ).

To draw a chemical potential energy diagram:

1. Write a balanced chemical equation with the enthalpy change (ΔH). The ΔH value must be in kilojoules (kJ), not molar enthalpy.
2. Label the axes: reaction coordinate (x-axis) and potential energy or enthalpy (y-axis).
3. Represent the reactants and products on the diagram.
    - **Exothermic Reaction:** Products have lower potential energy than reactants.
    - **Endothermic Reaction:** Products have higher potential energy than reactants.
4. Indicate the enthalpy change (ΔH) as the energy difference between reactants and products. An arrow pointing downwards represents a negative (exothermic) ΔH, while an arrow pointing upwards represents a positive (endothermic) ΔH.
5. Include the **activation energy barrier**, representing the energy required to start the reaction.

Example: Formation of Magnesium Oxide

Consider the reaction: $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

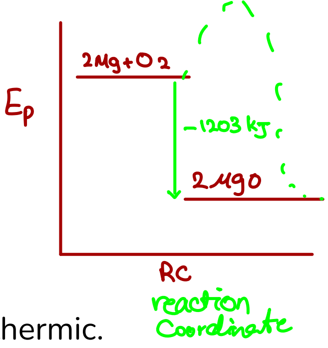
The molar enthalpy of formation of magnesium oxide (MgO) is $-601.6, \text{kJ/mol}$.

First, calculate the enthalpy change (ΔH) for the reaction:

$$\Delta H = \text{molar enthalpy} \times \text{moles} = -601.6, \text{kJ/mol} \times 2, \text{mol} = -1203, \text{kJ}$$

In the chemical potential energy diagram:

- Reactants: $2Mg(s) + O_2(g)$ at a certain energy level.
- Products: $2MgO(s)$ at a lower energy level since the reaction is exothermic.
- Enthalpy Change: A downward arrow indicating $-1203, \text{kJ}$.



It's important to note that if the stoichiometric coefficients of the equation are doubled, the enthalpy change must also be doubled.

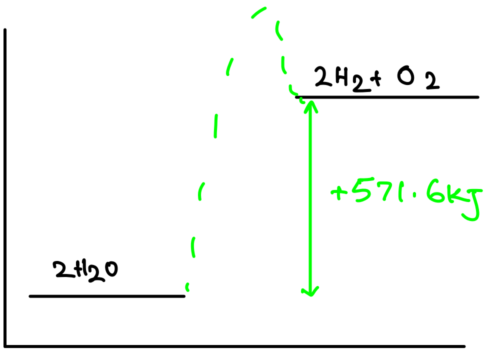
Practice Problem

Draw a chemical potential energy diagram for the decomposition of water (H_2O). You will need to determine the balanced equation, calculate ΔH from the given molar enthalpy, and then construct the diagram.

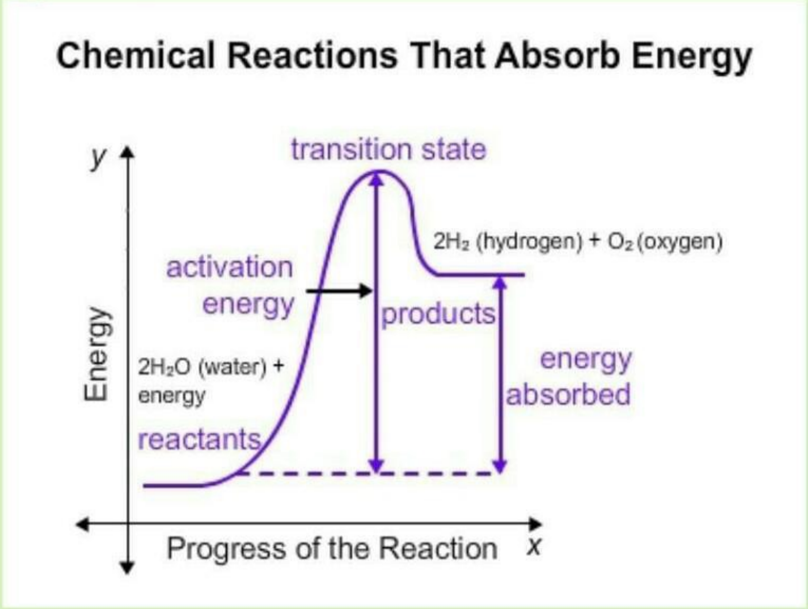
$\Delta_d H_m = 285.8 \text{ kJ/mol}$

$2H_2O(l) \longrightarrow 2H_2 + O_2$

$\Delta_d H = 2 \times 285.8 = 571.6 \text{ kJ}$



Explanation:


Chemical Reactions That Absorb Energy

Predicting Enthalpy Change of a Reaction

Recap of Previous Methods

- **Molar enthalpy** is the central concept for linking different methods of communicating enthalpy or sets of data.
- If you have an **equation with a ΔH value**, the equation provides the number of moles of each chemical and the associated ΔH value, allowing you to determine molar enthalpy.
- If the equation has an **energy term**, a positive value indicates energy as a reactant, while a negative value indicates energy released as a product.
- Molar enthalpy can be used to find a value in a **chemical potential energy diagram**, where the change in potential energy (equated to ΔH) is plotted against the reaction coordinate.
- **Calorimetry experiments** allow you to solve for the heat released or absorbed, and knowing the moles of the limiting reagent allows you to solve for molar enthalpy.

Predicting ΔH for Equations

- The new skill is predicting ΔH for equations, which allows you to predict molar enthalpy.
- This is useful when you don't have calorimetry data but have data tables for the reactants and products (e.g., combustion of propane).
- This technique helps answer real-world questions like how much energy is released when 10 grams of propane are combusted.

Enthalpies of Formation

- Before using enthalpies of formation to solve for ΔH, we must understand what enthalpies of formation are.

- **Formation reactions** show the formation of a compound or molecule from its elements in their elemental states.

  Example: Formation of water: $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$

  - The change in enthalpy for this reaction is $-285.6kJ$.

- **Formation tables** tabulate these values specifically for the product.

- A relative scale is set where every atom in its elemental form has a molar enthalpy value of zero.

  - For hydrogen gas: $H_2(g) = 0$
  - For oxygen gas: $O_2(g) = 0$
  - For liquid water: $\Delta H = -285.6kJ$

- The values are assigned to the product (e.g., water).

- This is analogous to the Celsius scale, which is arbitrary compared to the Kelvin scale.

- This allows us to predict any reaction type.

Predicting ΔH for Any Reaction

- You can predict ΔH for any reaction using formation values (single replacements, double replacements, and combustion reactions).
- To predict the change in enthalpy for a reaction, use the following equation:

$$\Delta H = \Sigma(moles\ of\ product \times molar\ enthalpy\ of\ formation\ of\ product) - \Sigma(moles\ of\ reactant \times molar\ enthalpy\ of\ formation\ of\ \text{reactants)}$$

- Data tables are under standard conditions, and ΔH will also be under standard conditions.

Example: Predicting ΔH for a Reaction

- Consider the reaction: $CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(s)$
- To find the change in enthalpy:

$$\Delta H = [1\ mole\ Ca(OH)_2 \times (-986.1\ kJ/mol)] - [1\ mole\ CaO \times (-634.9\ kJ/mol) + 1\ mole\ H_2O \times (-285.8\ kJ/mol)]$$

- $\Delta H = -64.8\ kJ$

Example: Predicting ΔH for Combustion of Propane

- We want to solve for the standard enthalpy change for the reaction of propane and then use that to answer what if 10 grams of propane is burned.
- First, balance the chemical equation: $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$
  - Propane (C_3H_8)
  - Oxygen (O_2)
  - Carbon Dioxide (CO_2)
  - Water Vapor (H_2O)
- Important note: Burning anything in an open system yields gaseous water (water vapor), not liquid water, except in cellular processes where liquid water is produced.## Predicting Enthalpy Change

We can use **formation values** to predict the **enthalpy change** (ΔH) of a reaction without conducting a whole experiment.

Formula for Predicting Enthalpy Change

To predict the enthalpy change at standard conditions for a combustion reaction, we use the following formula:

Σ (moles × molar enthalpy of formation of products) - Σ (moles × molar enthalpy of formation of reactants)

- Products should be calculated before reactants.

Example: Combustion of Propane

Consider the combustion reaction of propane (C_3H_8) with oxygen (O_2) to form carbon dioxide (CO_2) and water (H_2O):

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$$

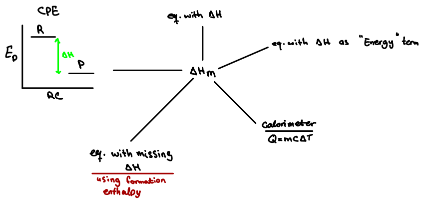
Calculation

1. **Products First:**
   - Carbon Dioxide (CO_2): 3 moles × -393.5 kJ/mol = -1180.5 kJ
   - Water Vapor (H_2O): 4 moles × -241.8 kJ/mol = -967.2 kJ
2. **Reactants:**
   - Propane (C_3H_8): 1 mole × -103.8 kJ/mol = -103.8 kJ
   - Oxygen (O_2): 5 moles × 0 kJ/mol = 0 kJ (since oxygen is in its elemental form)
3. **Enthalpy Change Calculation:**
   - $\Delta H_{combustion} = [(-1180.5) + (-967.2)] - [-103.8 + 0] = -2043.9\ \text{kJ}$

   The negative sign indicates that the reaction releases energy (exothermic).

Key Takeaway

For the balanced equation where one mole of propane reacts with five moles of oxygen, 2043.9 kJ of heat is released. The equation and the enthalpy term are linked; this value is specific to the reaction of one mole of propane.



Molar Enthalpy and Solving for Enthalpy Change

Molar Enthalpy Definition

Molar Enthalpy: A variation of molar enthalpy is change in enthalpy divided by moles (ΔH/moles).

Using Molar Enthalpy to Solve for ΔH

The change in enthalpy (ΔH) can be found using the formula:

$\Delta H =$ moles × molar enthalpy

Example Problem

What is the enthalpy change for the combustion of 10.00 grams of propane?

1. **Find the Molar Enthalpy of Propane Combustion.**
    - From the previous calculation, we know that when one mole of propane reacts, -2043.9 kJ of energy is released.
    - Therefore, the molar enthalpy of combustion of propane is -2043.9 kJ/mol.
2. **Calculate Moles of Propane**
    - Given: 10.00 grams of propane (C_3H_8).
    - Molar mass of propane: 44.11 g/mol.
    - Moles of propane = mass / molar mass = 10.00 g / 44.11 g/mol = 0.227 moles.
3. **Calculate Enthalpy Change**
    - $\Delta H =$ moles × molar enthalpy
    - $\Delta H = 0.227 \text{ moles} \times -2043.9 \text{ kJ/mol} = -463.96 \text{ kJ}$
    - Rounding to three significant figures: $\Delta H = -464 \text{ kJ}$

Therefore, 464 kJ of energy would be released for the combustion of 10.00 grams of propane.

Steps for Solving Multi-Step Problems

1. **Balance the equation.**
2. **Predict ΔH for the balanced equation.**
3. **Determine molar enthalpy.**
4. **Solve for the ΔH of the specific question.**

Key Consideration

It's crucial to remember that the first calculation of ΔH is for the entire equation (i.e., per mole), while the second calculation is for the specific amount given in the problem (e.g., 10 grams).

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Scholars Of Calgary Northwest

Chemistry 30, Unit 1: Thermochemical Changes, 10 Questions

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1.

the volume of the solution

15.00 g of calcium hydroxide, $Ca(OH)_{2(s)}$, is mixed with 700.0 mL of 0.500 mol/L $H_2SO_{4(aq)}$ in an insulated cup The temperature rises by 31.2°C. Assume that all of the calcium hydroxide reacts. The density of the final solution is 1.00 g/mL and a specific heat capacity of 4.184 J/g °C.

What is the heat of reaction per mol of $Ca(OH)_{2(s)}$?

A) -91380 J/mol
B) -260 kJ/mol
C) -452 kJ/mol
D) -1000 kJ/mol

Correct. Your answer=C, Correct answer=C

$$\text{mol } Ca(OH)_2 = 15g \times \frac{1\text{ mol}}{74g} = \boxed{0.202\text{ mol}}$$ Limiting re.

$$\text{mol } H_2SO_4 = 0.7 \times 0.5 = 0.35\text{ mol}$$

$$\text{mass} = 700\text{ mL} \times \frac{1g}{mL} = 700g$$

$$Q = mc\Delta T = 700 \times 4.184 \times 31.2 = 91378.56\text{ J}$$
$$= 91.4\text{ kJ} \rightarrow \text{for } \underline{\underline{0.202\text{ mol}}}$$

$$\Delta H = \frac{91.4}{0.202} = 452\text{ kJ/mol}$$

Explanation:

$$H_2SO_{4(aq)} + Ca(OH)_{2(s)} \longrightarrow \overset{CaSO_4}{\cancel{CaCl_{2(aq)}}} + H_2O_{(l)}$$

Since all of $Ca(OH)_{2(s)}$ is used up, it must be the limiting reagent.

Molar mass $Ca(OH)_{2(s)}$ = 74.10 g/mol

mol $Ca(OH)_{2(s)}$:
= 15.00 g/74.10 g/mol
= 0.202 mol

mass of reactants = 700.0 g

$Q = m \times c \times \Delta T$:
= 700.0g × 4.184 J/g °C 31.2°C
= 91379 J
= 91.4 kJ

$$\Delta H = \frac{91.4\text{ kJ}}{0.202\text{ mol}}$$
= 452 kJ/mol$Ca(OH)_{2(s)}$

2.

$$C_3H_{8(g)} + 5O_{2(g)} \longrightarrow 3CO_{2(g)} + 4H_2O_{(g)} + 2043.9\text{ kJ}$$

If 512 g of propane are burned in an automobile engine, the amount of energy released (to the nearest tenth) is __________ MJ.

Incorrect. Your answer=23.8, Correct answer=23.7

$$\text{mol } C_3H_8 = 512g \times \frac{1\text{ mol}}{44.11\ g} = 11.61$$

C: 12.01 × 3 = 36.03
H: 1.01 × 8 = 8.08
+ = 44.11

1 mol — 2043.9 kJ
11.61 mol — ?

heat = 11.61 × 2043.9 = 23724.2 kJ = 23.7 MJ

I took this as $(12.01\times3)+8 = 44.03\ \frac{g}{mol}$

Explanation:

$$\text{mol of } C_3H_8 = \frac{m}{M} = \frac{512g}{44.11\text{ g/mol}} = 11.61\text{ mol}$$

heat given off = 2043.9 kJ/mol × 11.61 mol = 23724 kJ = 23.7 MJ

3.

$$\boxed{2C_4H_{10(g)} + 13O_{2(g)} \longrightarrow 8CO_{2(g)} + 10H_2O_{(g)} + \text{heat}}$$

A lighter contains the fuel butane. How much heat is released when 0.500 g of butane in a disposable lighter is burned to form gaseous carbon dioxide and water vapour?

A) –2877.3 kJ
B) –2657.3 kJ
C) –125.7 kJ
D) –22.9 kJ

$$\text{mol } C_4H_{10} = 0.5g \times \frac{1\text{ mol}}{\underset{58.14}{\cancel{(12.01\times4)+(1.01\times10)}}\,g} =$$

$$= 0.00859993\ldots\text{ mol}$$

$$\Delta_cH = [(10\times-241.8)+(8\times-393.5)]-[2\times-125.7]$$

$$= -5314.6\ \text{kJ}/2\text{mol } C_4H_{10} = -2657.3\ \text{kJ}/\text{mol } C_4H_{10}$$

1 mol : −2657.3 kJ
0.0086 ?

Heat = −22.85 kJ = −22.9 kJ

***Correct**. Your answer=D, Correct answer=D*

Explanation:

$\Delta_c H$:
$= [10\Delta_f H\,(H_2O_{(g)}) + 8\Delta_f H(CO_{2(g)})] - [2\Delta_f H\,(C_4H_{10(g)})]$
= [10mol ×(-241.8 k/mol) + 8mol ×(-393.5 kJ/mol)] – [2mol ×(-125.7 kJ/mol)]
= -5314.6 kJ/2 mol butane
= -2657.3 kJ/mol molar mass C_4H_{10}

mol of butane burned:

$$= \frac{n}{M}$$

$$= \frac{0.500g}{58.14\text{ g/mol}}$$

= 0.0085999 mol

heat released:
= 0.0085999 mol × -2657.3 kJ/mol
= -22.9 kJ

4. The molar heat of solution for $NaOH_{(s)}$ is –44.5 kJ/mol. If 40.0 g of $NaOH_{(s)}$ is dissolved in 200.0 mL of water in a calorimeter, the temperature of the water would increase by

→ m = 200 g

A) 53.1°C
B) 2124°C
C) 5.31×10^{-2}°C
D) 26.6°C

$$\text{mols NaOH} = 40g \times \frac{1\text{ mol}}{\underset{40}{\cancel{22.99+1.01+16}}\quad g} = 1\text{mol NaOH}$$

$Q_w = mc\Delta T$

$$-44.5\times1000\,\cancel{J} = 200\,\cancel{g} \times 4.186\,\frac{\cancel{J}}{\cancel{g}\,°C} \qquad \Delta T \Rightarrow$$

$$\Delta T = \frac{-44500}{837.2} = 53.1\,°C$$

***Incorrect**. Your answer=C, Correct answer=A*

Explanation:

mol NaOH:

$$= \frac{m}{M}$$

$$= \frac{40.0\text{ g}}{40.0\text{ g/mol}}$$

= 1.00 mol

heat gained by water:

= 44.5 kJ
= 44,500 J

I used the kJ, forgot to Convert it to J!

$Q = m \times c \times \Delta T$
44,500 J = 200.0 g × 4.184 J/g°C × ΔT
ΔT = 53.1 °C

5.

Given the reaction:
$Cl_2O_{(g)} + 3OF_{2(g)} \longrightarrow 2ClF_{3(l)} + 2O_{2(g)} + 394\text{ kJ}$

→ released → exothermic → $\Delta H_r < 0$ → $\Delta H_r = -394\text{kJ}$

What is the heat of reaction, ΔH_r, for the reaction
$\frac{2}{3}ClF_{3(l)} + \frac{2}{3}O_{2(g)} \longrightarrow \frac{1}{3}Cl_2O_{(g)} + OF_{2(g)}$ → reverse of what we see in the question → Endothermic → $\Delta H_r > 0$

÷ by 3

A) +394 kJ
B) -131 kJ
C) -1182 kJ
D) +131 kJ

reverse: $\frac{2ClF_3}{3} + \frac{2O_2}{3} + \frac{394\text{ kJ}}{3} \longrightarrow \frac{Cl_2O}{3} + \frac{3OF_2}{3}$

$\Delta H_r = +\frac{394}{3} = +131\text{ kJ}$

Correct. *Your answer=D, Correct answer=D*

Explanation:

The given equation is exothermic and for this reaction ΔH_r = -394 kJ.
The reaction $\frac{2}{3}ClF_{3(l)} + \frac{2}{3}O_{2(g)} \longrightarrow \frac{1}{3}Cl_2O_{(g)} + OF_{2(g)}$ is the reverse of the given equation divided by 3.

Heat of reaction for this reaction:
= $\frac{1}{3}$ × + 394 kJ
= +131 kJ

6.

For the reaction

$HA_{(aq)} + BOH_{(aq)} \longrightarrow BA_{(aq)} + H_2O_{(l)}$

volume = 200 mL => mass = 200 g

the temperature rises by 0.320°C when 100.0 mL of 0.0500 mol/L $HA_{(aq)}$ and 100.0 mL of 0.0500 mol/L $BOH_{(aq)}$ are mixed in a calorimeter. The specific heat capacity of the solutions is 4.184 J/g °C and the density of the solutions is 1.00 g/mL.

From this experimental data, which answer is correct?

A) $HA_{(aq)} + BOH_{(aq)} \longrightarrow BA_{(aq)} + H_2O_{(l)} + 26.8\text{ kJ}$
B) $HA_{(aq)} + BOH_{(aq)} \longrightarrow BA_{(aq)} + H_2O_{(l)} + 53.6\text{ kJ}$
C) $H^+_{(aq)} + OH^-_{(aq)} \longrightarrow H_2O_{(l)} + 268\text{ kJ}$
D) $HA_{(aq)} + BOH_{(aq)} \longrightarrow BA_{(aq)} + H_2O_{(l)}\ \Delta H = -5.36\text{ kJ}$

$Q = mc\Delta T$
$= 200\text{ g} \times 4.184 \frac{J}{g°C} \times 0.32°C$
$= 267.8\text{ J} = 0.2678\text{ kJ}$

mol HA = mol BOH = 0.1 × 0.05 = 0.005

0.005 mol   0.2678 kJ
1 mol   53.56 kJ

Correct. *Your answer=B, Correct answer=B*

Explanation:

Q = $m \times c \times \Delta T$
= 200.0 g × 4.184 J/g°C × 0.320 °C
= 267.8 J
= 0.2678 kJ

mol $HA_{(aq)}$:
= mol $BOH_{(aq)}$
= $c \times V$
= 0.100 L x 0.0500 mol/L

= 0.00500 mol

heat released per mol of reactant:

$$= \frac{0.2678\ \text{kJ}}{0.00500\ \text{mol}}$$

= 53.5 kJ/mol

7. When a $\overset{m}{\underline{\text{50.0 g}}}$ sample of an unknown metal is heated from $\overset{\Delta T}{\underline{\text{–10.0°C to 60.0°C}}}$, $\overset{Q}{\text{452 J}}$ of energy is absorbed. The specific heat capacity of the metal is

$452 = 50 \times C \times 70 \Rightarrow C = 0.129\ \text{J/g°C}$

   A) 0.129 J/g·°C
   B) 0.181 J/g·°C
   C) 9.04 J/g·°C
   D) 0.113 J/g·°C

Correct. *Your answer=A, Correct answer=A*

Explanation:

heat gained:
$Q = m \times c \times \Delta T$
452 J = 50.0 g × c × 70.0 °C
c = 0.129 J/g °C

8.

Reaction 1:

$2Fe_{(s)} + \frac{3}{2}O_{2(g)} \longrightarrow Fe_2O_{3(s)}$
ΔH = -825 kJ

Reaction 2:

$C(graphite) + O_{2(g)} \longrightarrow CO_{2(g)}$
ΔH = -394 kJ

$\Delta H_r = [3 \times -394] - [2 \times -825] = +468\ \text{kJ}$

What is the enthalpy change for the reaction:
$\underset{-825\text{kJ}}{2Fe_2O_{3(s)}} + \underset{0}{3\ C_{(s)}} \longrightarrow \underset{-394\text{kJ}}{3CO_{2(g)}} + \underset{0}{4Fe_{(s)}}$

   A) -1219 kJ
   B) +468 kJ
   C) +357 kJ
   D) -431kJ

Correct. *Your answer=B, Correct answer=B*

Explanation:

Reaction 1 (× -2):

$2Fe_2O_{3(s)} \longrightarrow 4Fe_{(s)} + 3O_{2(g)}$
ΔH = (-2)(-825 kJ)
= +1650 kJ

Reaction 2 (× 3):

$3C(graphite) + 3O_{2(g)} \longrightarrow 3CO_{2(g)}$
ΔH = 3(-394 kJ)
= -1182 kJ

Add:

$2Fe_2O_{3(s)} + 3C(graphite) \longrightarrow 4Fe_{(s)} + 3CO_{2(g)}$
$\Delta H = +468$ kJ

9.

$C_5H_{12(l)} + 8O_{2(g)} \longrightarrow 5CO_{2(g)} + 6H_2O_{(g)}$

$\Delta H_{comb} = -3244.8$ kJ/mol of $C_5H_{12(l)}$
$\Delta H^\circ_f(H_2O_{(g)}) = -241.8$ kJ/mol
$\Delta H^\circ_f(CO_{2(g)}) = -393.5$ kJ/mol

$\Delta H_r = \Delta H_p - \Delta H_R$
$-3244.8 = ((6\times -241.8) + (5\times -393.5)) - X \Rightarrow$
$x = -173.5$ kJ/mol

Use this information to calculate the heat of formation, $\overset{x}{\underline{\Delta H^\circ_f}}$, for pentane, $C_5H_{12(l)}$.

A) -173.5 kJ/mol
B) -6663.1 kJ/mol
C) -325.2 kJ/mol
D) -94.9 kJ/mol

Correct. Your answer=A, Correct answer=A

Explanation:

$\Delta H_{comb} = [5\Delta H^\circ_f(CO_{2(g)}) + 6\Delta H^\circ_f(H_2O_{(g)})] - [\Delta H^\circ_f (C_5H_{12(l)})]$
$-3244.8\text{ kJ} = [(5\text{ mol} \times -393.5\text{ kJ/mol} + 6\text{ mol} \times -241.8\text{ kJ/mol})] - [\Delta H^\circ_f (C_5H_{12(l)})]$
$\Delta H^\circ_f (C_5H_{12(l)}) = -173.5$ kJ/mol

10.

The decomposition of sulfur trioxide is represented by the equation

$2SO_{3(g)} \longrightarrow 2SO_{2(g)} + O_{2(g)}$
$\Delta H^\circ = +198$ kJ

$2SO_3 + 198\text{ kJ} \longrightarrow 2SO_2 + O_2 \Rightarrow$
$\frac{2SO_2}{2} + \frac{O_2}{2} \longrightarrow \frac{2SO_3}{2} + \frac{198\text{ kJ}}{2} \Rightarrow$
$SO_2 + \frac{1}{2}O_2 \longrightarrow SO_3 + 99\text{ kJ}$

Based upon this information, which one of the following thermochemical equations is correct?

A) $SO_{3(g)} \longrightarrow SO_{2(g)} + \frac{1}{2}O_{2(g)} + 198$ kJ
B) $SO_{2(g)} + \frac{1}{2}O_{2(g)} + 198\text{ kJ} \longrightarrow SO_{3(g)}$
C) $SO_{2(g)} + \frac{1}{2}O_{2(g)} \longrightarrow SO_{3(g)} + 99$ kJ
D) $SO_{3(g)} \longrightarrow SO_{2(g)} + \frac{1}{2}O_{2(g)} + 99$ kJ

$\Delta H > 0 \rightarrow$ endothermic $\Rightarrow$ the opposite reaction will be exothermic.

Incorrect. Your answer=D, Correct answer=C

Explanation:

Since $2SO_{3(g)} \longrightarrow 2SO_{2(g)} + O_{2(g)}$
$\Delta H^\circ = +198$ kJ

then, $2SO_{3(g)} + 198\text{ kJ} \longrightarrow 2SO_{2(g)} + O_{2(g)}$

or $2SO_{2(g)} + O_{2(g)} \longrightarrow 2SO_{3(g)} + 198$ kJ

divide through by 2:
$SO_{2(g)} + \frac{1}{2}O_{2(g)} \longrightarrow SO_{3(g)} + 99$ kJ

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