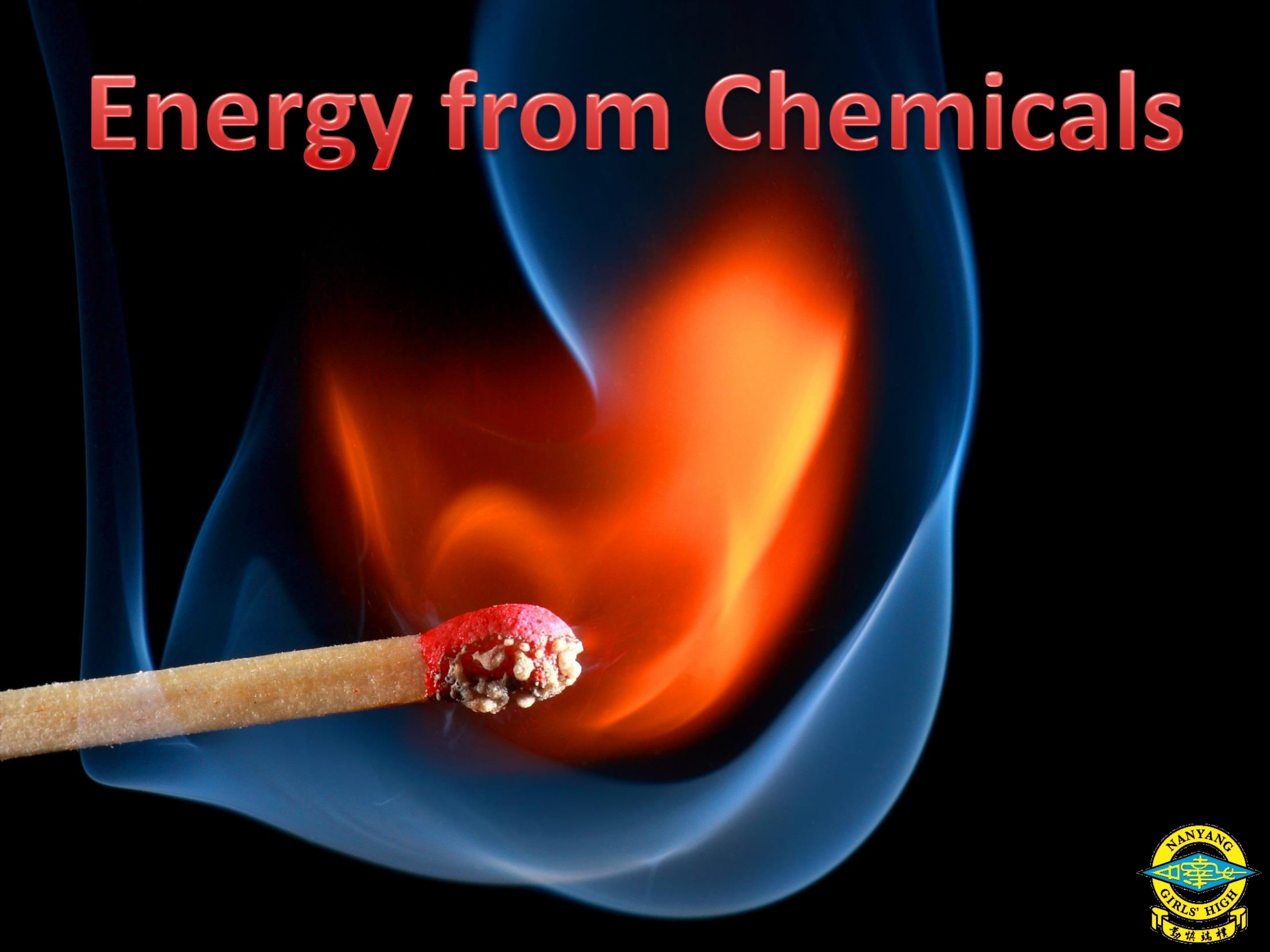
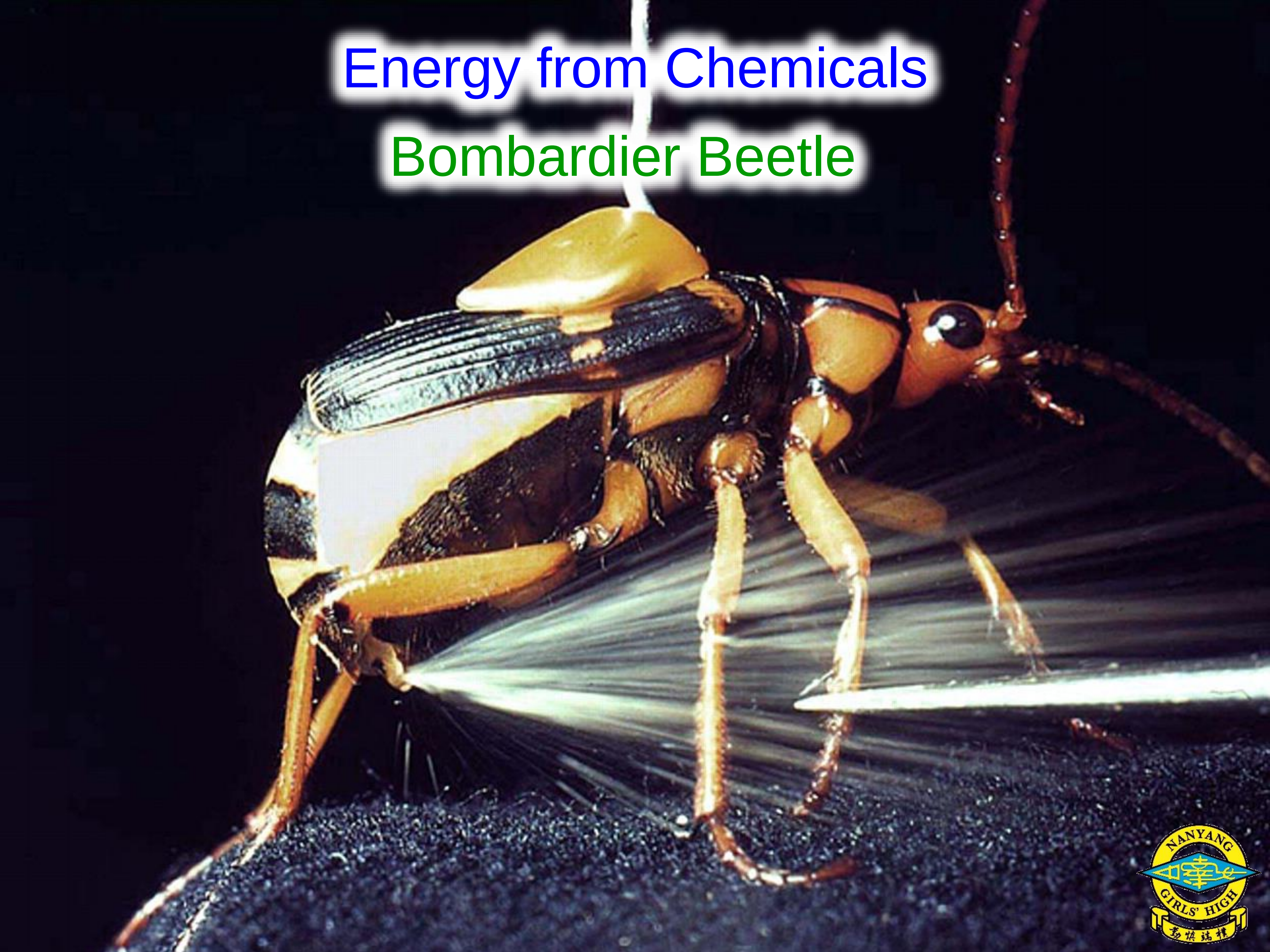
Energy from Chemicals



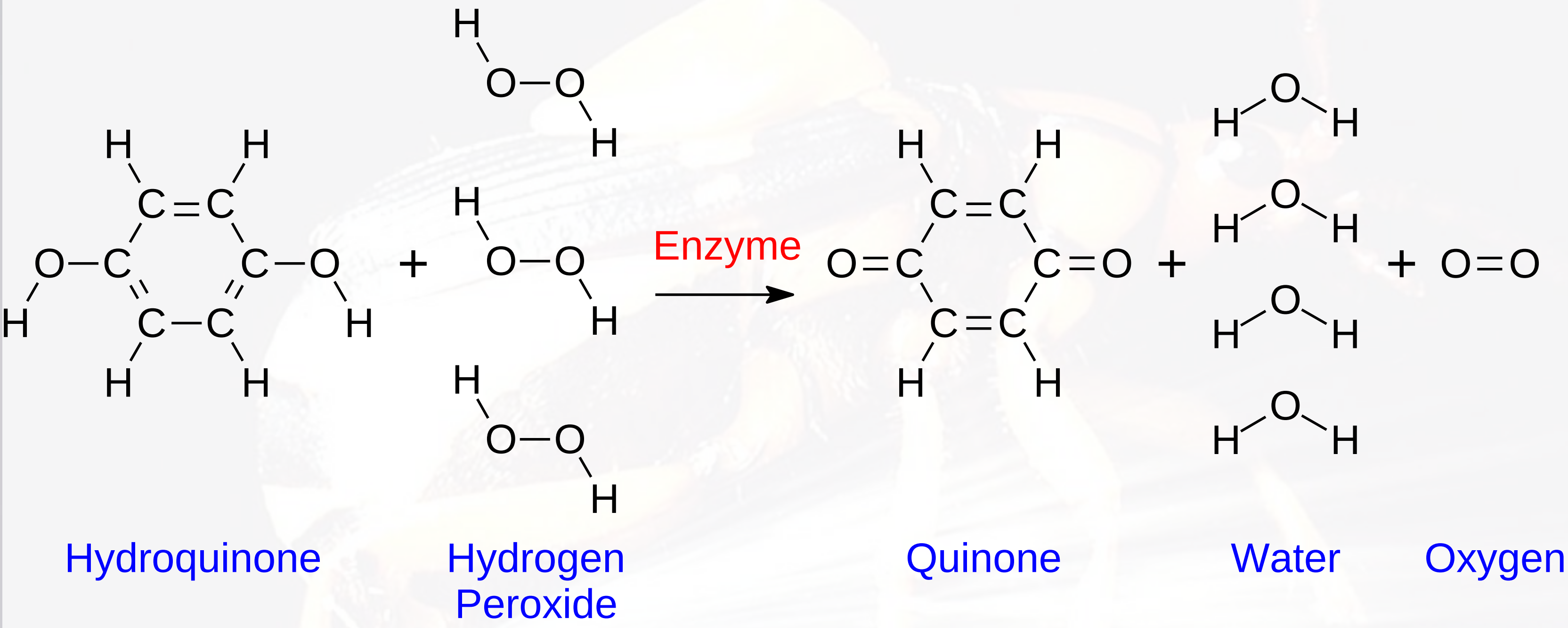
Energy from Chemicals

Bombardier Beetle



Energy from Chemicals

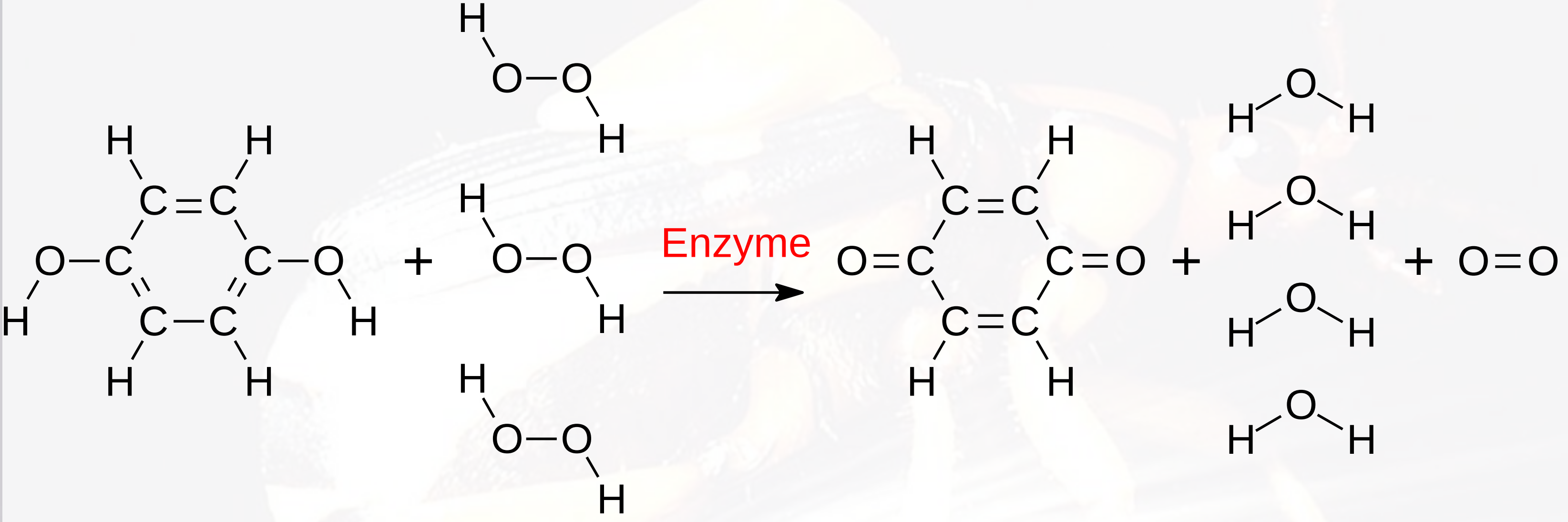
Bombardier Beetle



- This is an *exothermic* reaction that *releases* approximately *400 kJ of energy* for every mole of hydroquinone that reacts.

Energy from Chemicals

Bombardier Beetle



- Result? Whatever is attacking the bombardier beetle is sprayed with a *hot* and *corrosive liquid*.

Energy from Chemicals

Chemical reactions are always accompanied by a *change* in energy.

- What are some generalisations about *change*?

Energy from Chemicals

- Some examples of generalisations that can be made:

- **Change:**

→ Change is inevitable (unavoidable).

→ Change can have positive consequences or negative consequences.

→ Change can be reversible or irreversible.

→ Change can be steady, cyclic, random or chaotic.

Energy from Chemicals

- **Enduring Understandings:**
- Chemical reactions (change) can be spontaneous or non-spontaneous.
- Spontaneous reactions (changes) release energy which can be used to do work. Non-spontaneous reactions (changes) require energy in order to take place.

Energy from Chemicals

- **Essential Questions:**
- How are matter and energy related?
- Why are some chemical reactions exothermic and others endothermic?
- Why do some chemical reactions occur spontaneously, while others do not?
- Where does the heat and light that is given off by an exothermic reaction originate from, and where does it go to?

Energy from Chemicals

What must I know and understand about *energy changes* in Chemistry?

Energy from Chemicals

Learning Outcomes

Candidates should be able to:

a) Describe the meaning of enthalpy change in terms of exothermic (ΔH negative) and endothermic (ΔH positive) reactions.

b) Represent energy changes by energy profile diagrams, including reaction enthalpy changes and activation energies.

c) Describe bond breaking as an endothermic process and bond making as an exothermic process.

d) Explain overall enthalpy changes in terms of the energy changes associated with the breaking and making of covalent bonds.

e) Describe hydrogen, derived from water or hydrocarbons, as a potential fuel, reacting with oxygen to generate electricity directly in a fuel cell (details of the construction and operation of a fuel cell are not required).

- Singapore Examinations and Assessment
- Board University of Cambridge International Examinations
- Ministry of Education Singapore

Energy from Chemicals

What are the different types of energy changes that can take place in Chemistry?

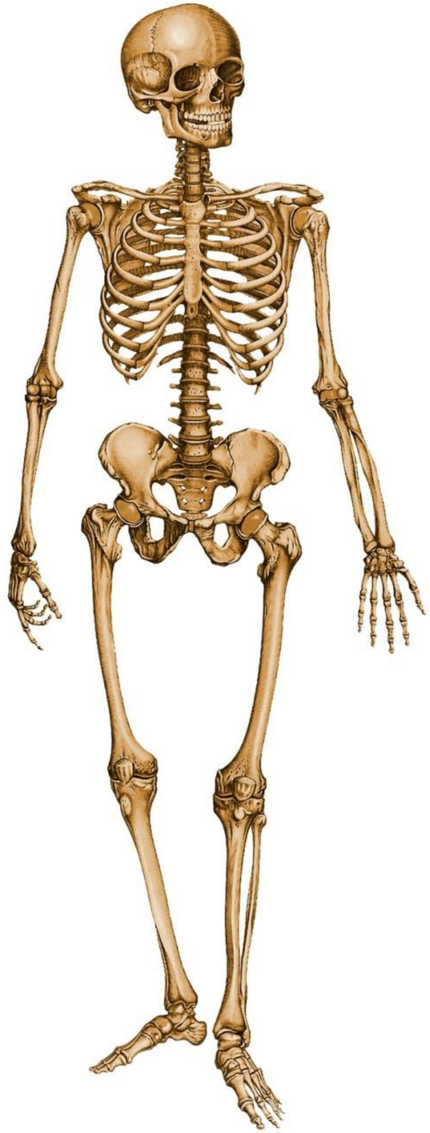
- **Exothermic**: Releases energy into the surroundings.
- **Endothermic**: Absorbs energy from the surroundings.

Energy from Chemicals – Exothermic

Exo is Greek for “outside.”
Thermasi is Greek for “to heat.”
The literal meaning of exothermic is therefore “to heat outside.”

Exo is a common term used throughout science. For example, insects are described as having an exoskeleton, which literally means “outside skeleton.” This accurately describes the animal’s hard outer casing.

Energy from Chemicals – Endothermic



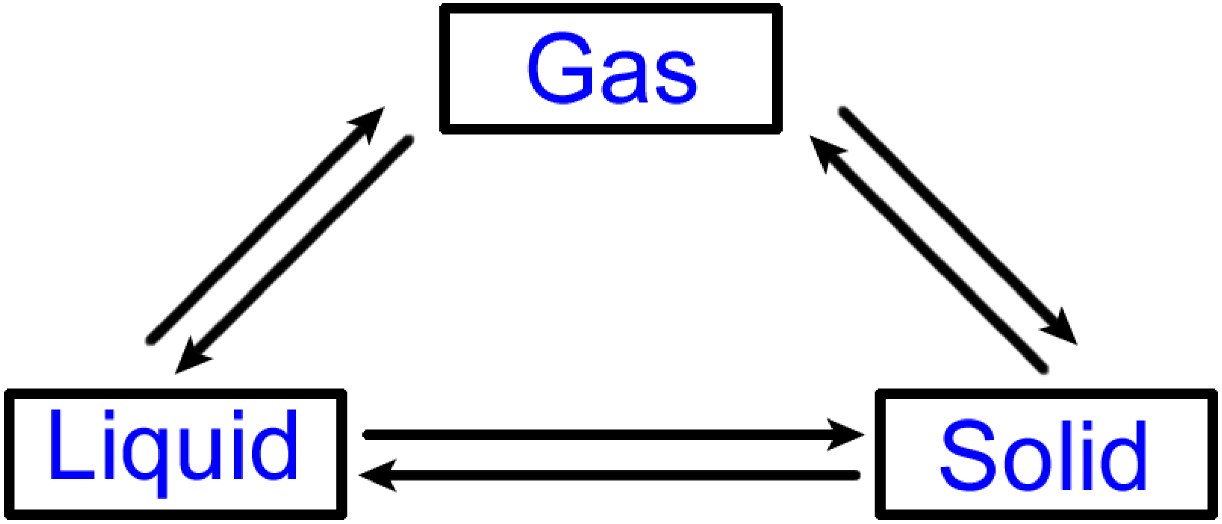
Endo is Greek for “inside.”
Thermasi is Greek for “to heat.”
The literal meaning of endothermic is therefore “to heat inside.”

Endo is a common term used throughout science. For example, mammals are described as having an endoskeleton, which literally means “inside skeleton.”

Energy from Chemicals

I remember from *kinetic particle theory* that changes in state also involve energy changes.

Energy from Chemicals

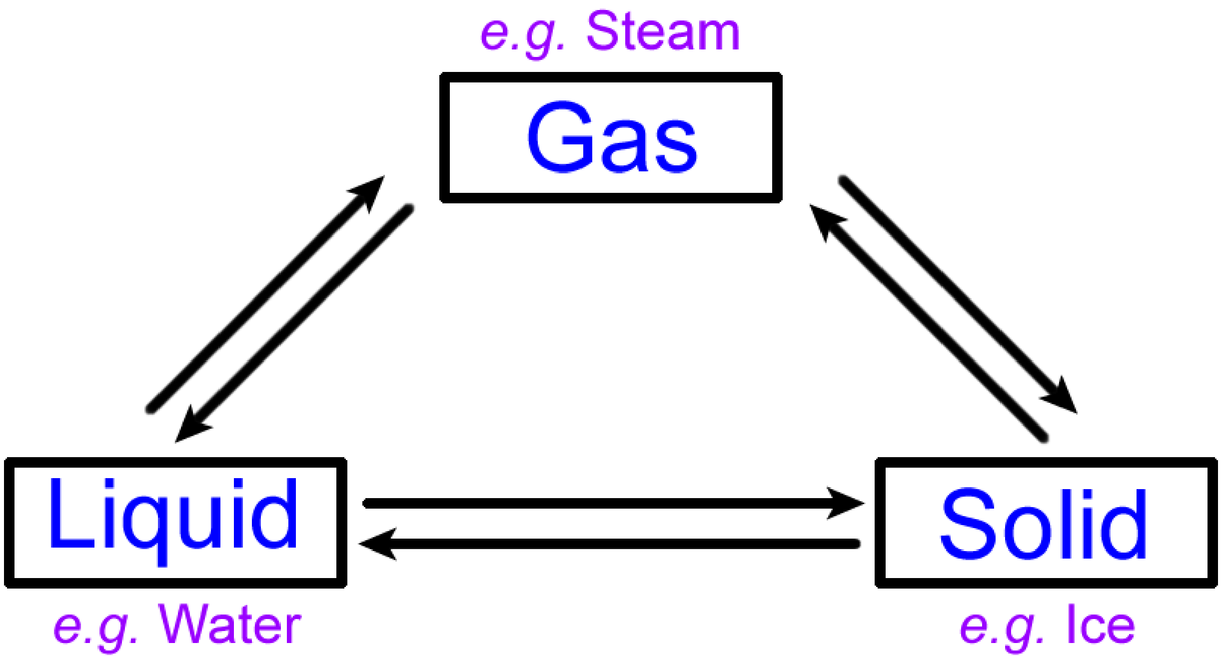


Gas

Liquid

Solid

Energy from Chemicals



e.g. Steam

Gas

Liquid

e.g. Water

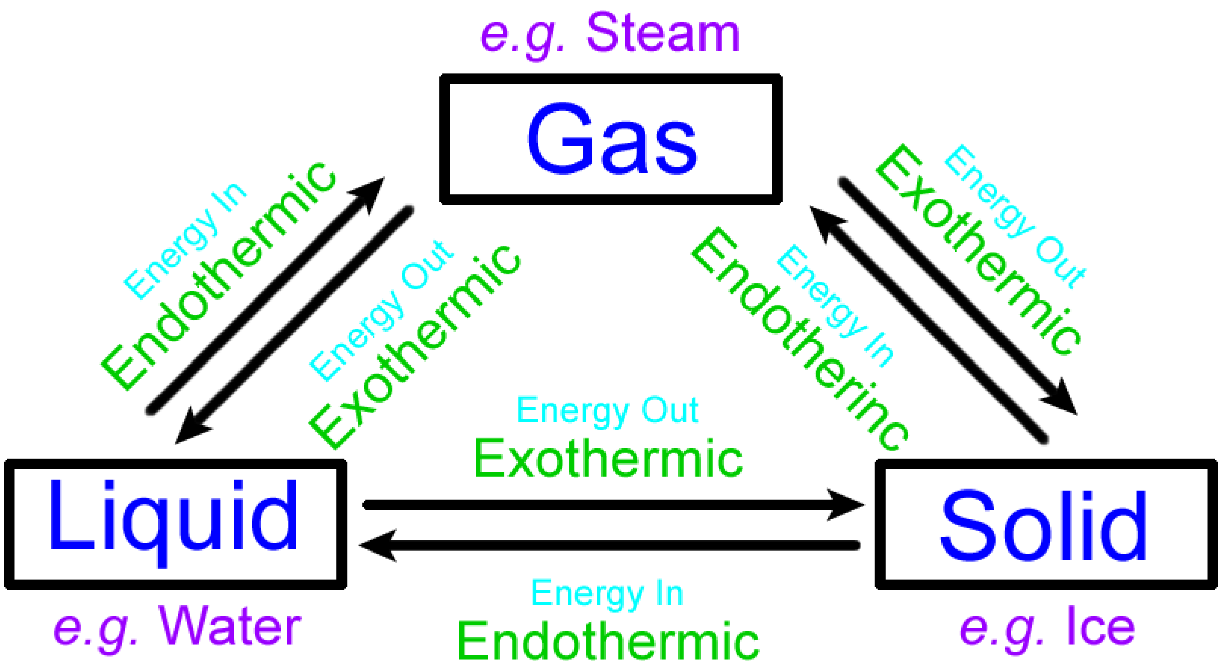
Solid

e.g. Ice

Remember:

energy $H_2O_{(g)}$ > energy $H_2O_{(l)}$ > energy $H_2O_{(s)}$

Energy from Chemicals



e.g. Steam

Gas

Energy In
Endothermic

Energy Out
Exothermic

Energy In
Endotherinc

Energy Out
Exothermic

Liquid

e.g. Water

Energy Out
Exothermic

Energy In
Endothermic

Solid

e.g. Ice

Remember:

energy $H_2O_{(g)}$ > energy $H_2O_{(l)}$ > energy $H_2O_{(s)}$

Energy from Chemicals

Phase changes involve a change in energy content of the system. Alcohol feels cold on the skin because the vaporisation of a liquid is an ***exothermic / endothermic*** process. As the liquid alcohol evaporates to form alcohol vapour, its energy content ***increases / decreases*** *i.e.* energy flows ***into / out of*** the system. Energy is ***taken from / given to*** the surroundings (the immediate surroundings is the skin), whose temperature therefore ***increases / decreases***.

Steam at 100 °C is said to cause a worse burn than boiling water at the same temperature because the condensation of a vapour to a liquid is an ***exothermic / endothermic*** process and energy flows ***into / out*** *of* the system ***from / into*** the surroundings. Once again the immediate surroundings include the skin whose temperature is therefore ***decreased / increased***.

Energy from Chemicals

Phase changes involve a change in energy content of the system. Alcohol feels cold on the skin because the vaporisation of a liquid is an ~~***exothermic***~~ **/** ***endothermic*** process. As the liquid alcohol evaporates to form alcohol vapour, its energy content ***increases*** **/** ~~***decreases***~~ *i.e.* energy flows ***into*** **/** ~~***out of***~~ the system. Energy is ***taken from*** **/** ~~***given to***~~ the surroundings (the immediate surroundings is the skin), whose temperature therefore ~~***increases***~~ **/** ***decreases***.

Steam at 100 °C is said to cause a worse burn than boiling water at the same temperature because the condensation of a vapour to a liquid is an ***exothermic*** **/** ~~***endothermic***~~ process and energy flows ~~***into***~~ **/** ***out*** *of* the system ~~***from***~~ **/** ***into*** the surroundings. Once again the immediate surroundings include the skin whose temperature is therefore ~~***decreased***~~ **/** ***increased***.

Energy from Chemicals

What about chemical reactions?

Exothermic Reactions

Energy from Chemicals

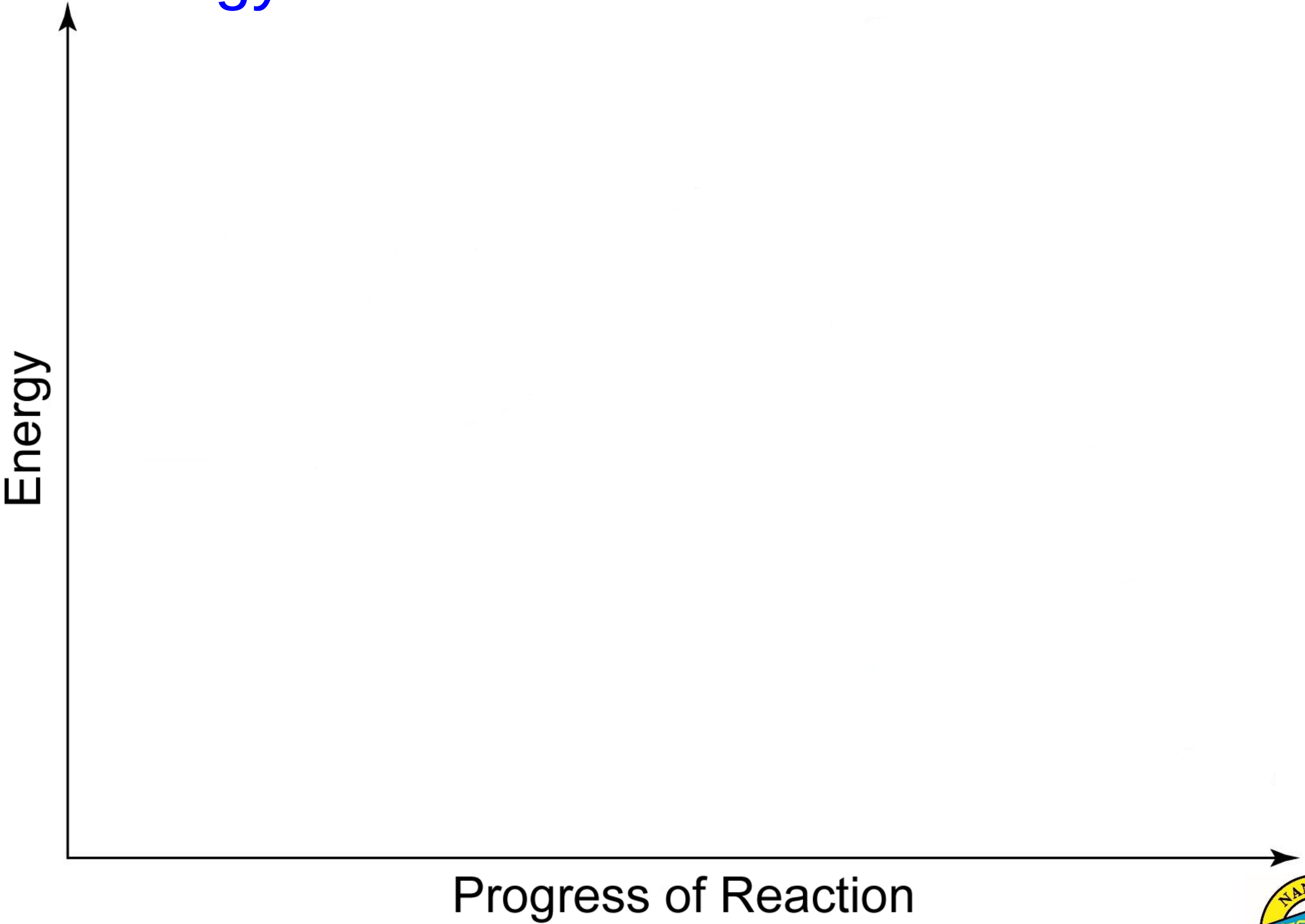
Important Things to Note…

- ΔH means change in enthalpy.

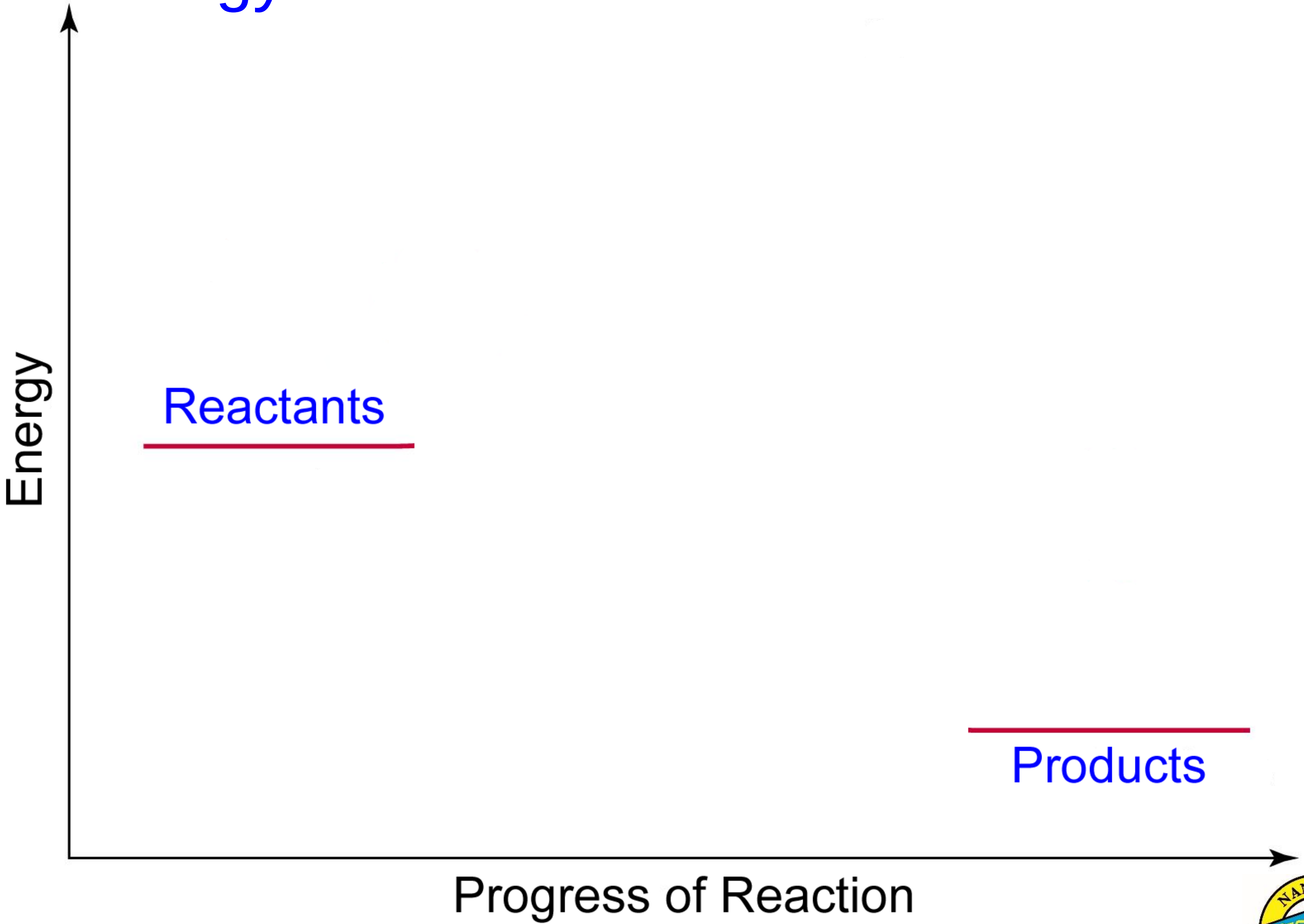
- An exothermic change releases energy into the surroundings. Numerical values of ΔH are negative.

- During an exothermic reaction, the temperature of the surroundings increases.

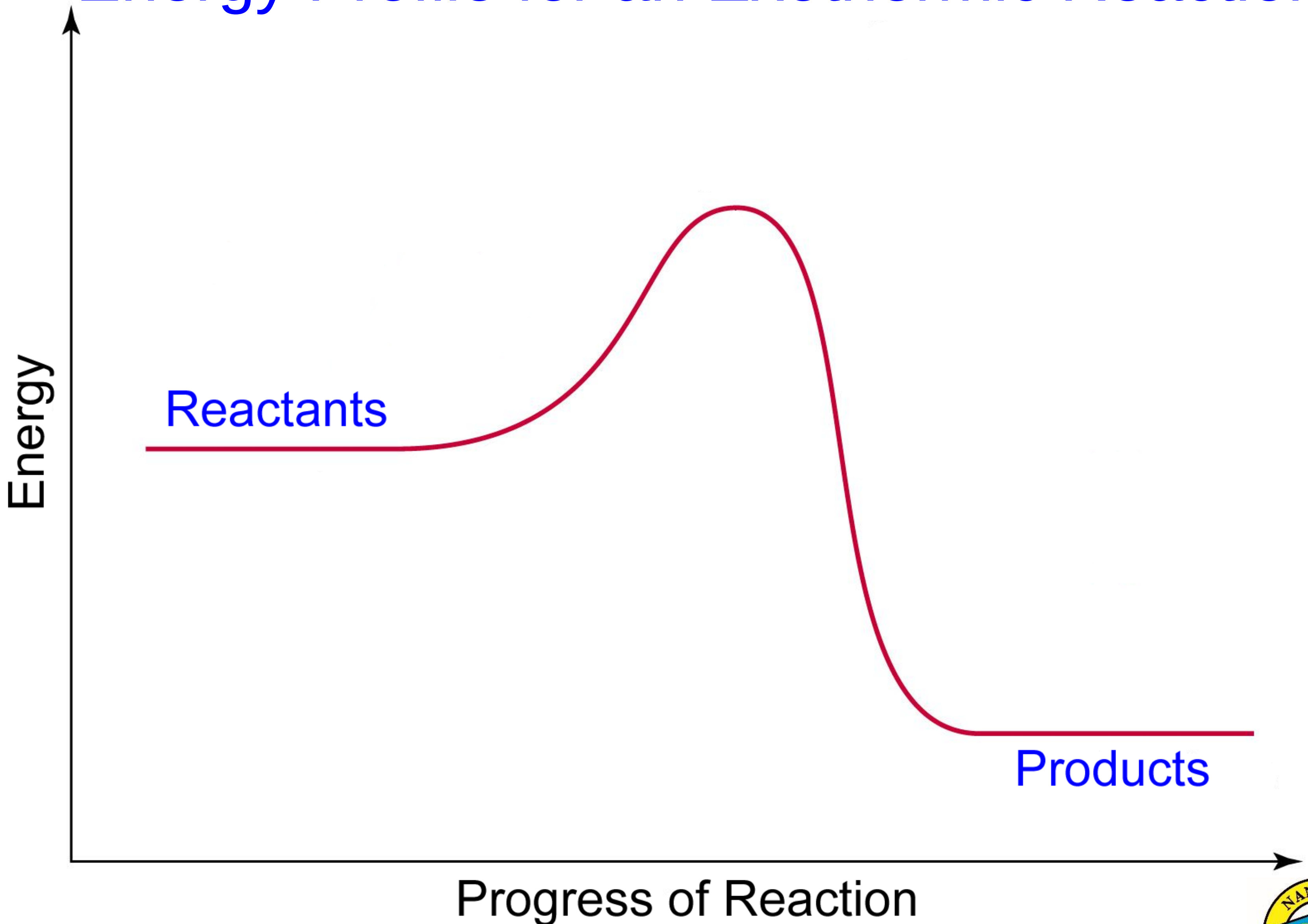
Energy Profile for an Exothermic Reaction



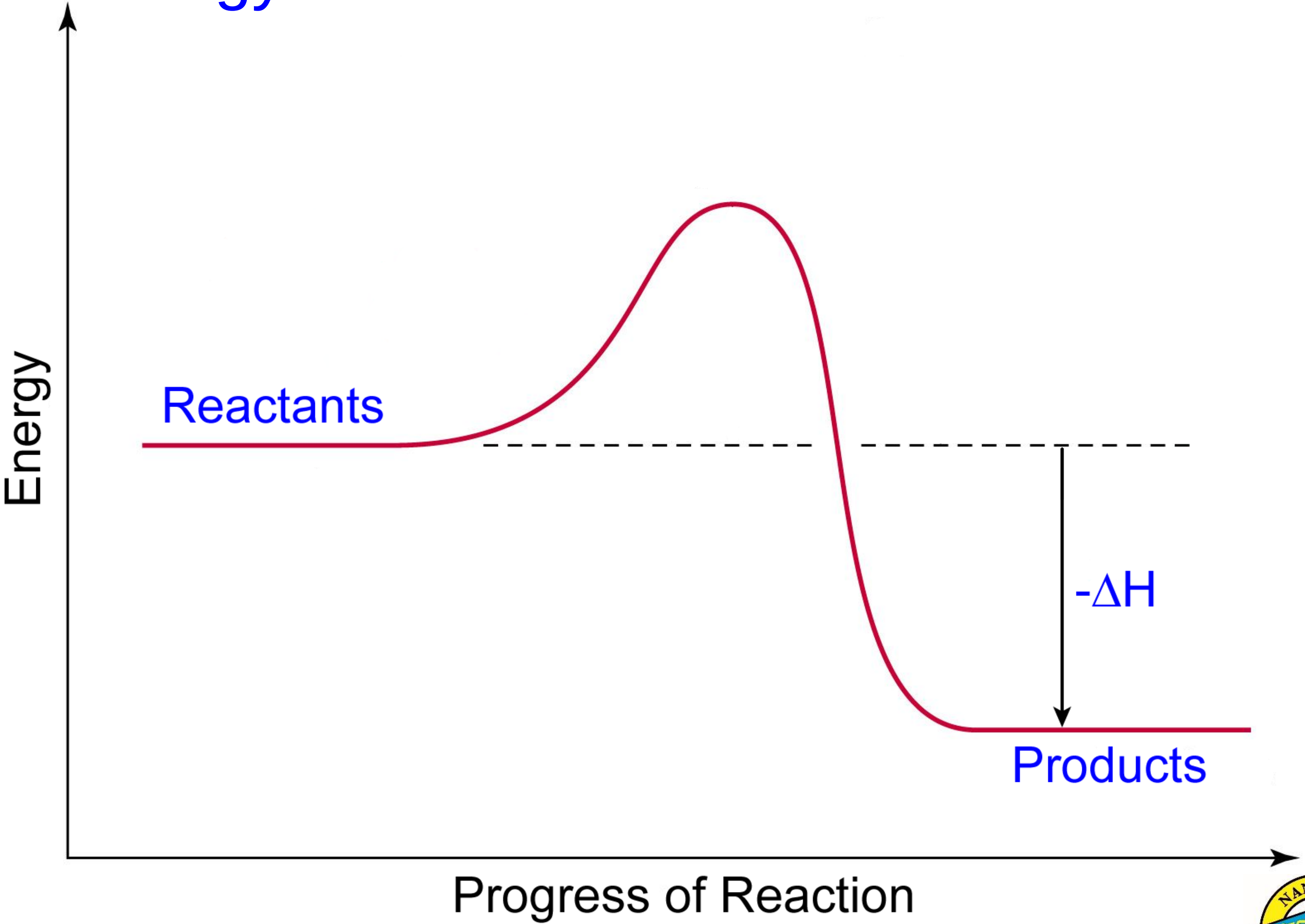
Energy Profile for an Exothermic Reaction



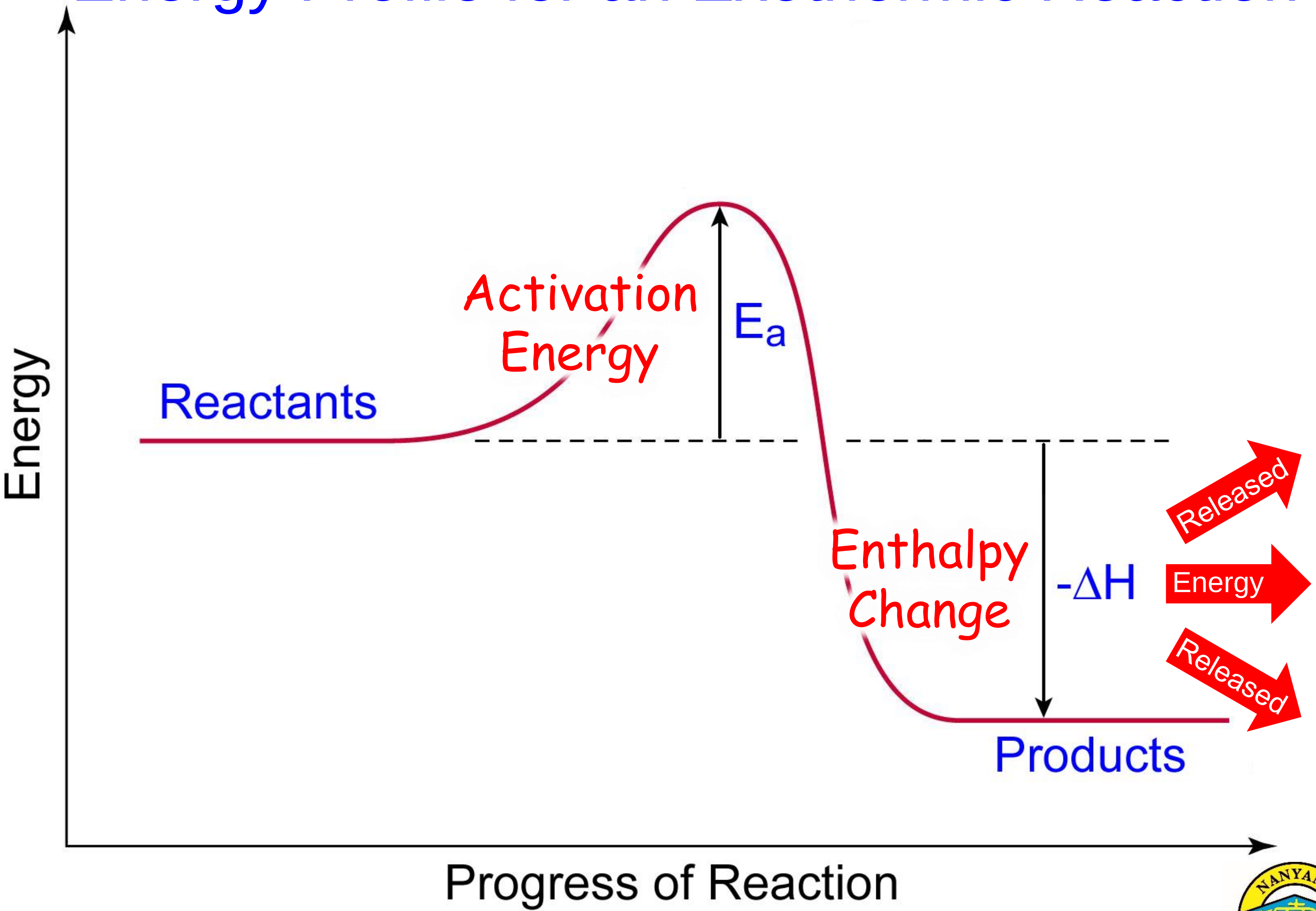
Energy Profile for an Exothermic Reaction



Energy Profile for an Exothermic Reaction



Energy Profile for an Exothermic Reaction



Energy

Reactants

Activation Energy

E_a

Enthalpy Change

$-\Delta H$

Released

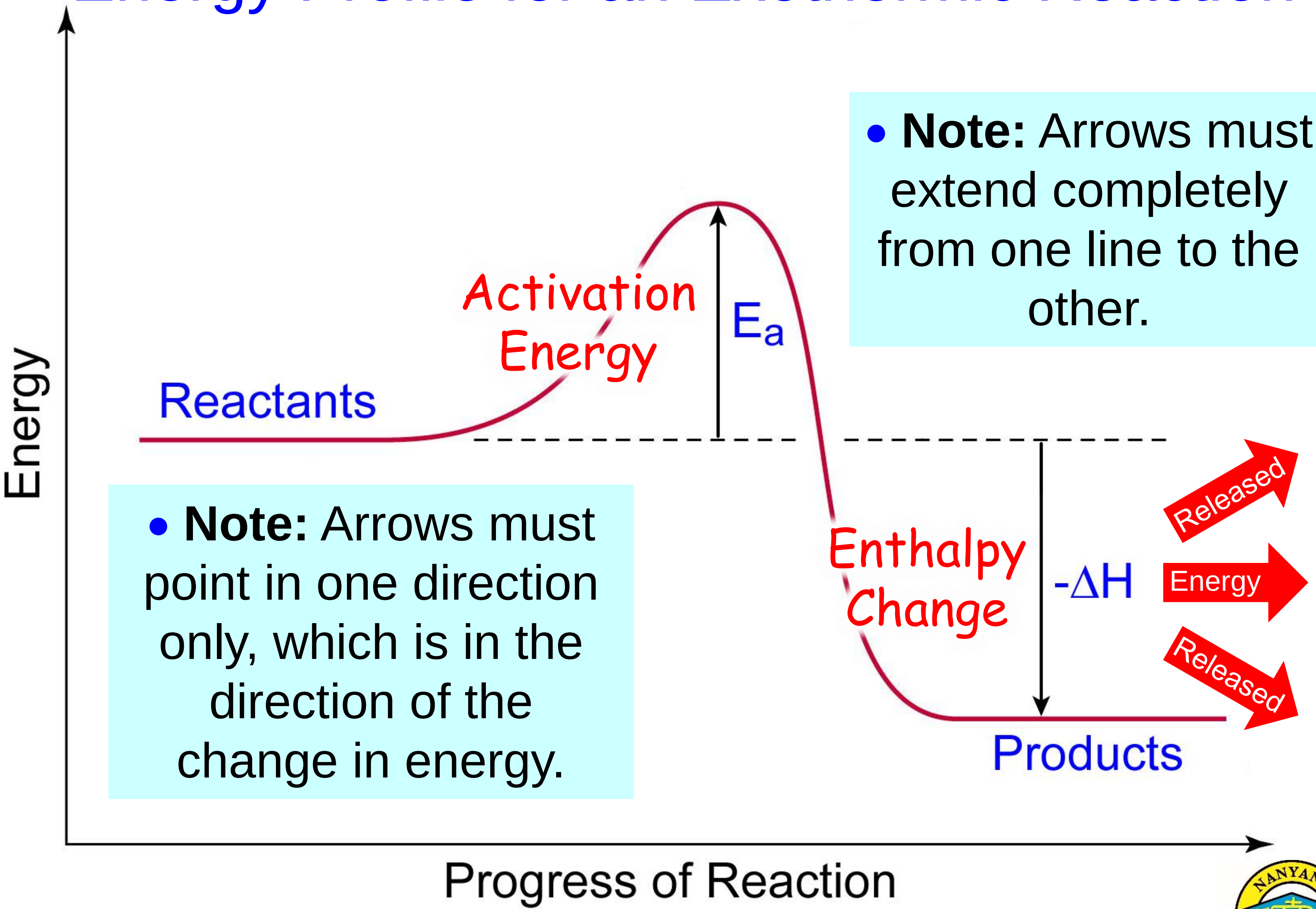
Energy

Released

Products

Progress of Reaction

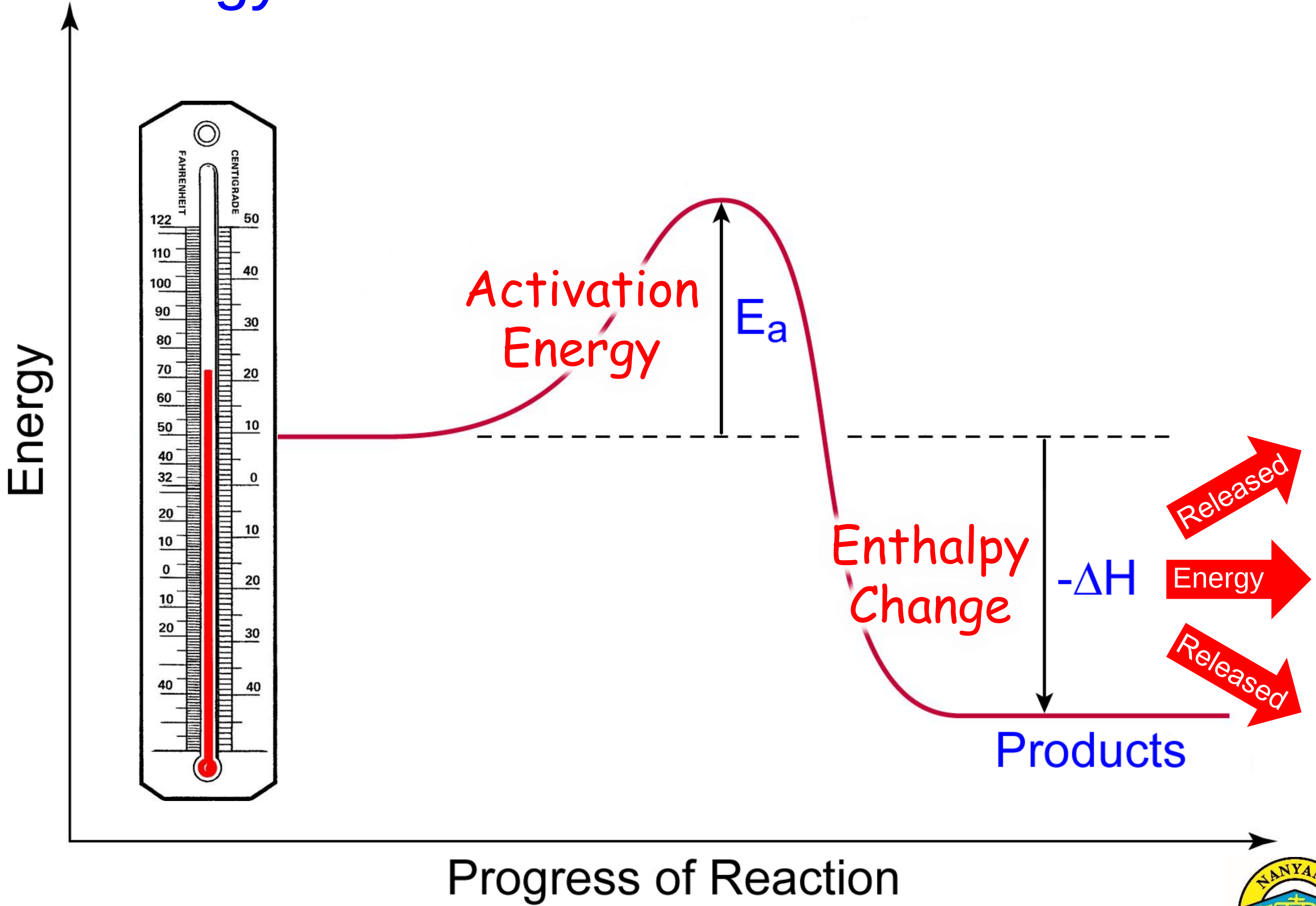
Energy Profile for an Exothermic Reaction



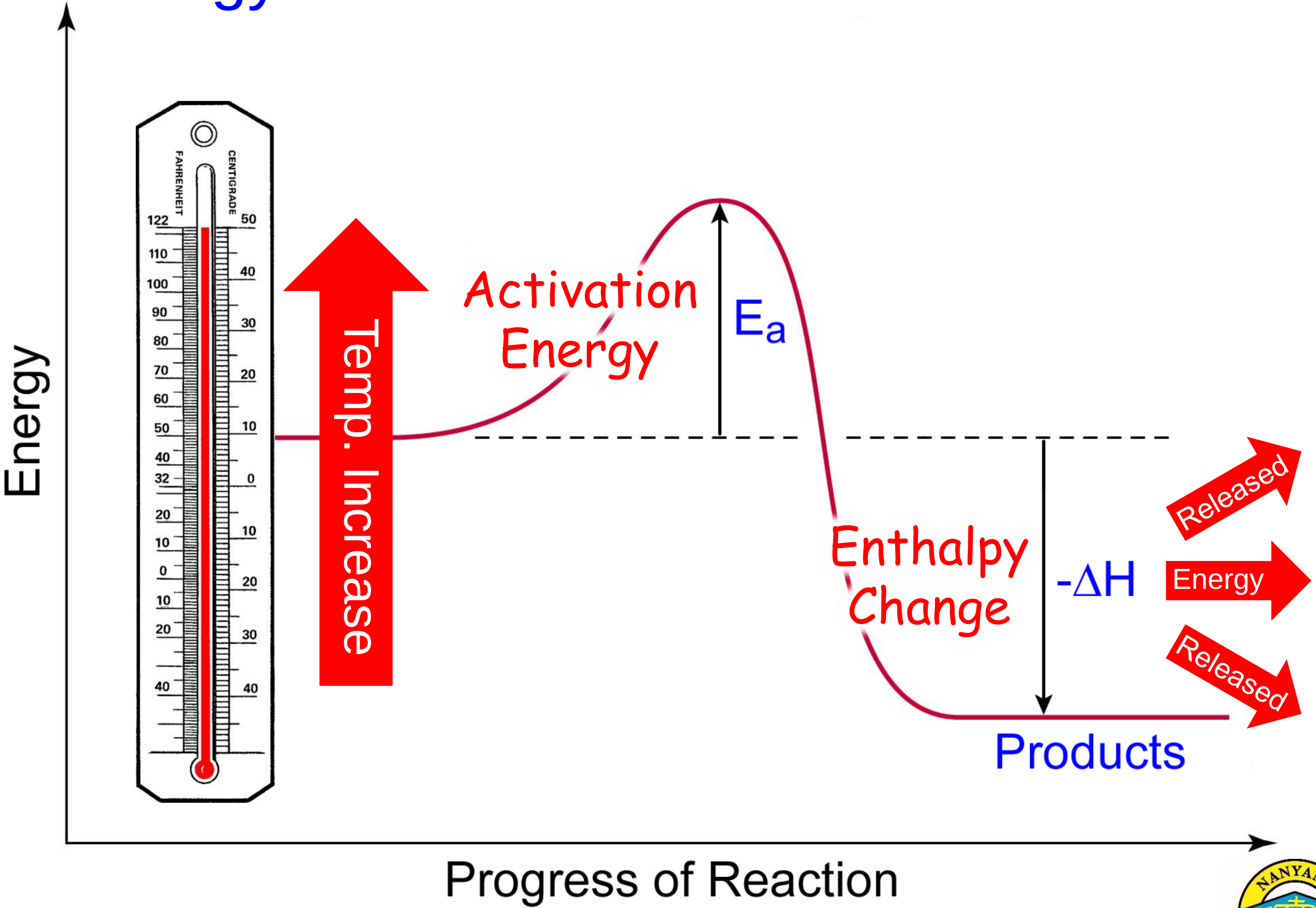
- **Note:** Arrows must extend completely from one line to the other.

- **Note:** Arrows must point in one direction only, which is in the direction of the change in energy.

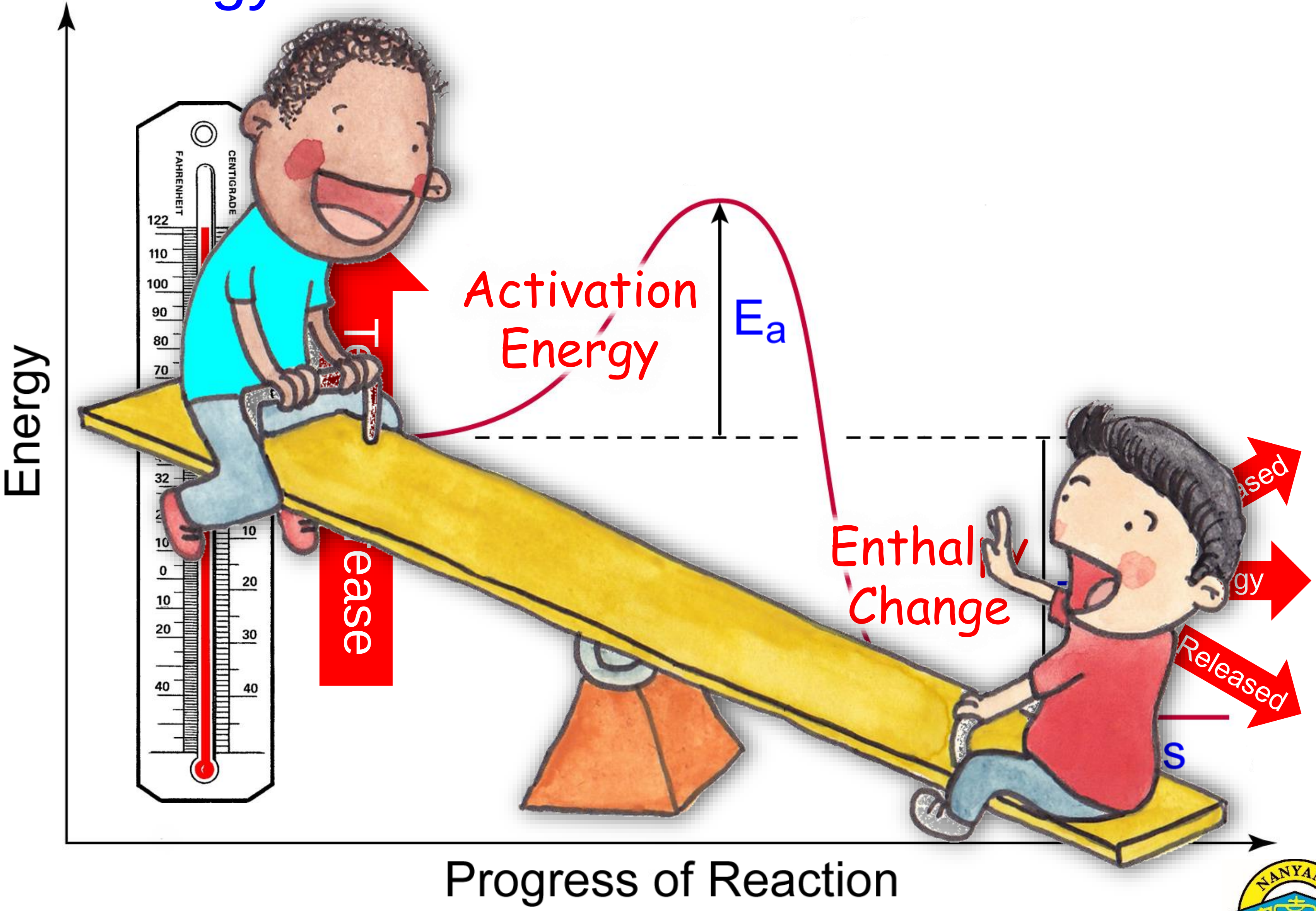
Energy Profile for an Exothermic Reaction



Energy Profile for an Exothermic Reaction



Energy Profile for an Exothermic Reaction



Energy from Chemicals

- It is very important to remember that the thermometer is part of the *surroundings* and *not* part of the *reaction*!

- As a consequence, the thermometer measures the *temperature* (or *energy*) of the *surroundings* and not the *temperature* (or *energy*) of the *reaction*!

Energy from Chemicals

- The particles in a reaction do not all react at once.
- People running a marathon run at different speeds and cross the finishing line at different times.
- It is the same for the particles in a reaction. The particles move at different speeds, with different amounts of kinetic energy, and will collide and react at different times.

Energy from Chemicals

What is *activation energy*?

Activation Energy

- Activation energy is the minimum amount of energy that must be supplied to a chemical in order for a reaction to take place.

Activation Energy

- For example, the phosphorus and sulfur in a match head will not react with the oxygen in the air until the match is struck against the sandpaper on the side of the match box.

Activation Energy

- Friction between the match head and sandpaper generates heat energy, which is used to break chemical bonds in the molecules of phosphorus, sulfur and oxygen, thus allowing them to react.

Exothermic Reactions

- All combustion reactions (where a fuel reacts with oxygen) are exothermic.

- For example, the combustion of methane:

$$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(l)}$$

$$\Delta H = -890 \text{ kJ / mol}$$

- For example, the combustion of ethanol:

$$C_2H_5OH_{(g)} + 3O_{2(g)} \rightarrow 2CO_{2(g)} + 3H_2O_{(l)}$$

$$\Delta H = -1371 \text{ kJ / mol}$$

Endothermic Reaction

Energy from Chemicals

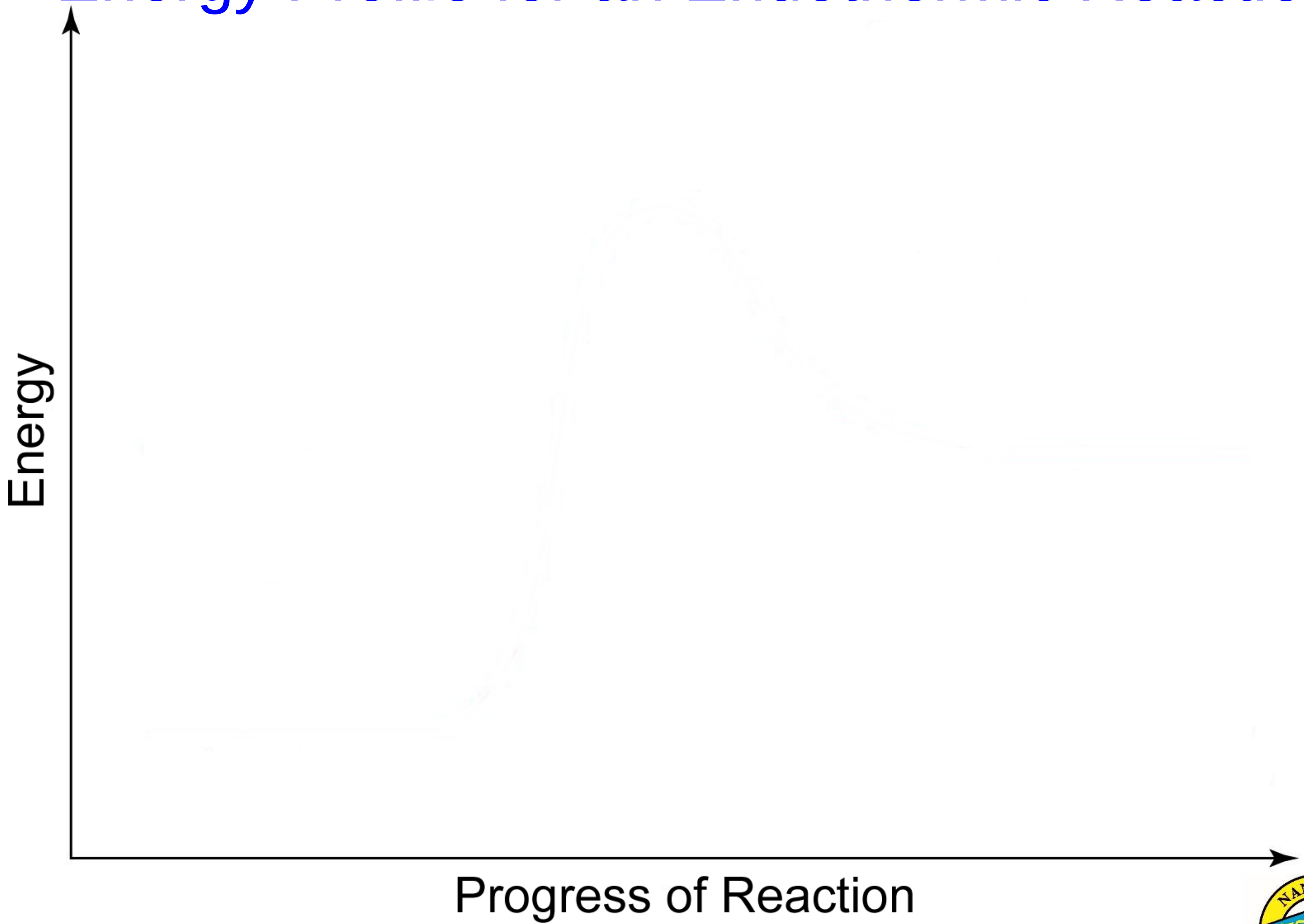
Important Things to Note…

- ΔH means change in enthalpy.

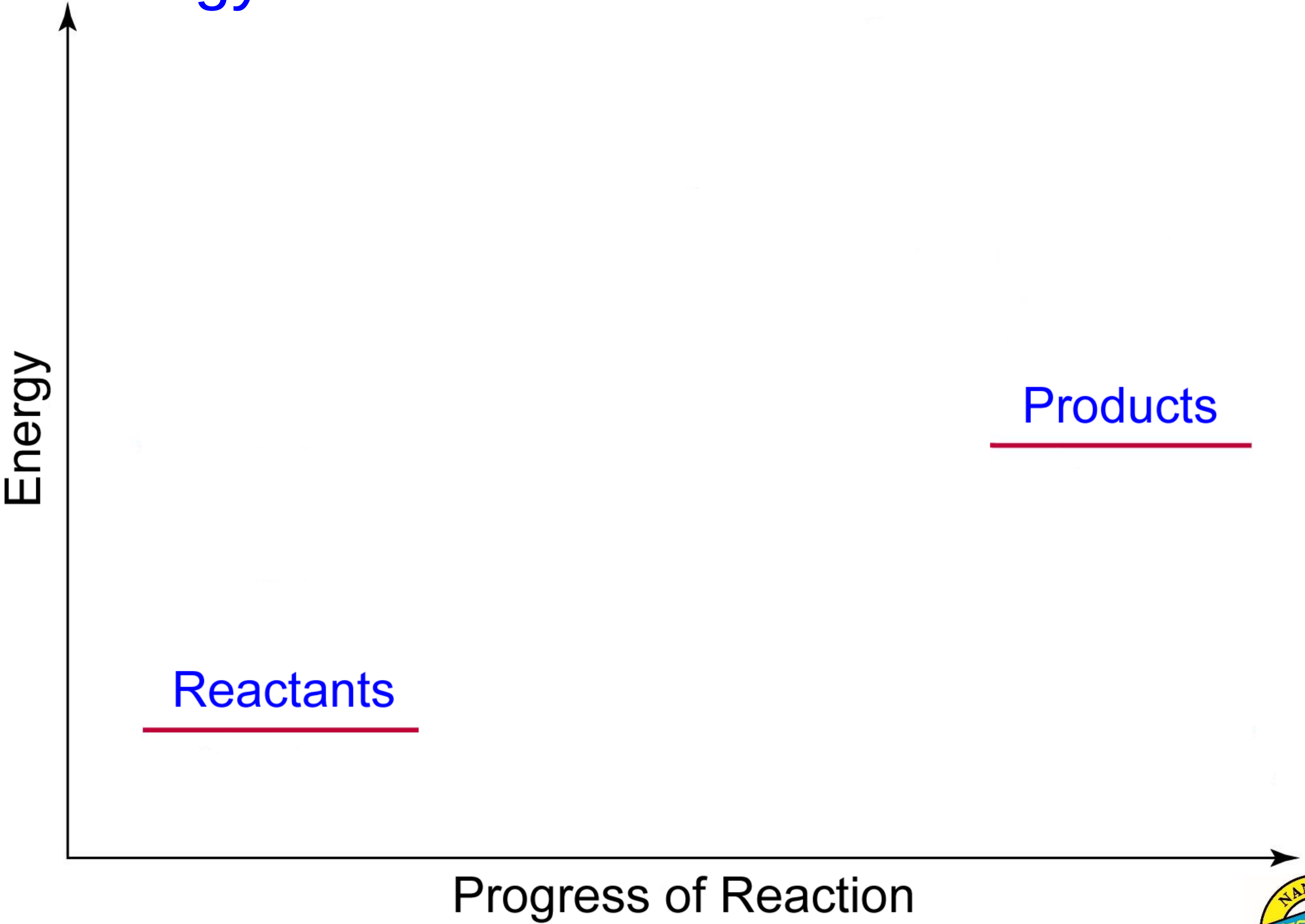
- An endothermic change absorbs energy from the surroundings. Numerical values of ΔH are positive.

- During an endothermic reaction, the temperature of the surroundings decreases.

Energy Profile for an Endothermic Reaction



Energy Profile for an Endothermic Reaction



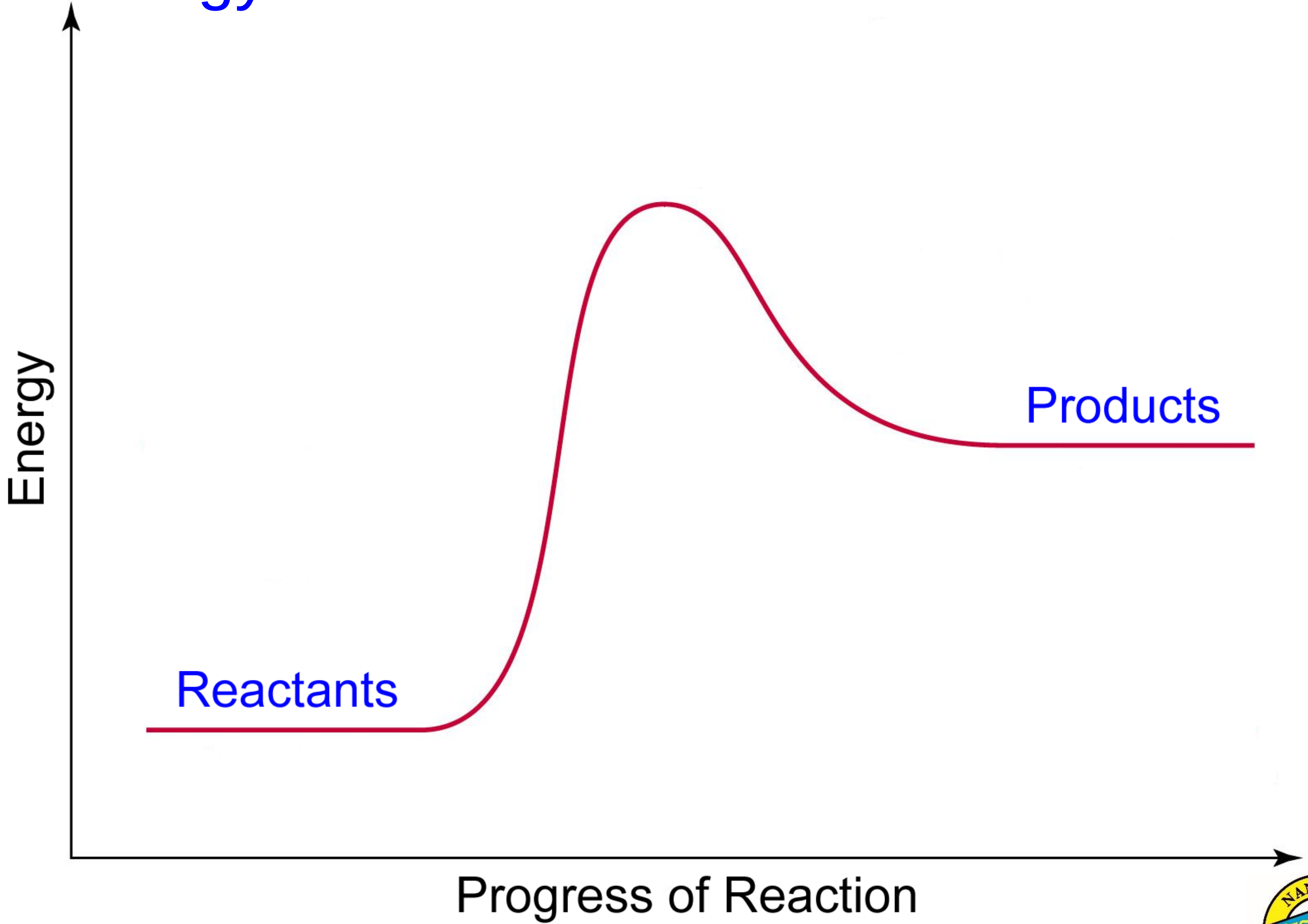
Energy

Products

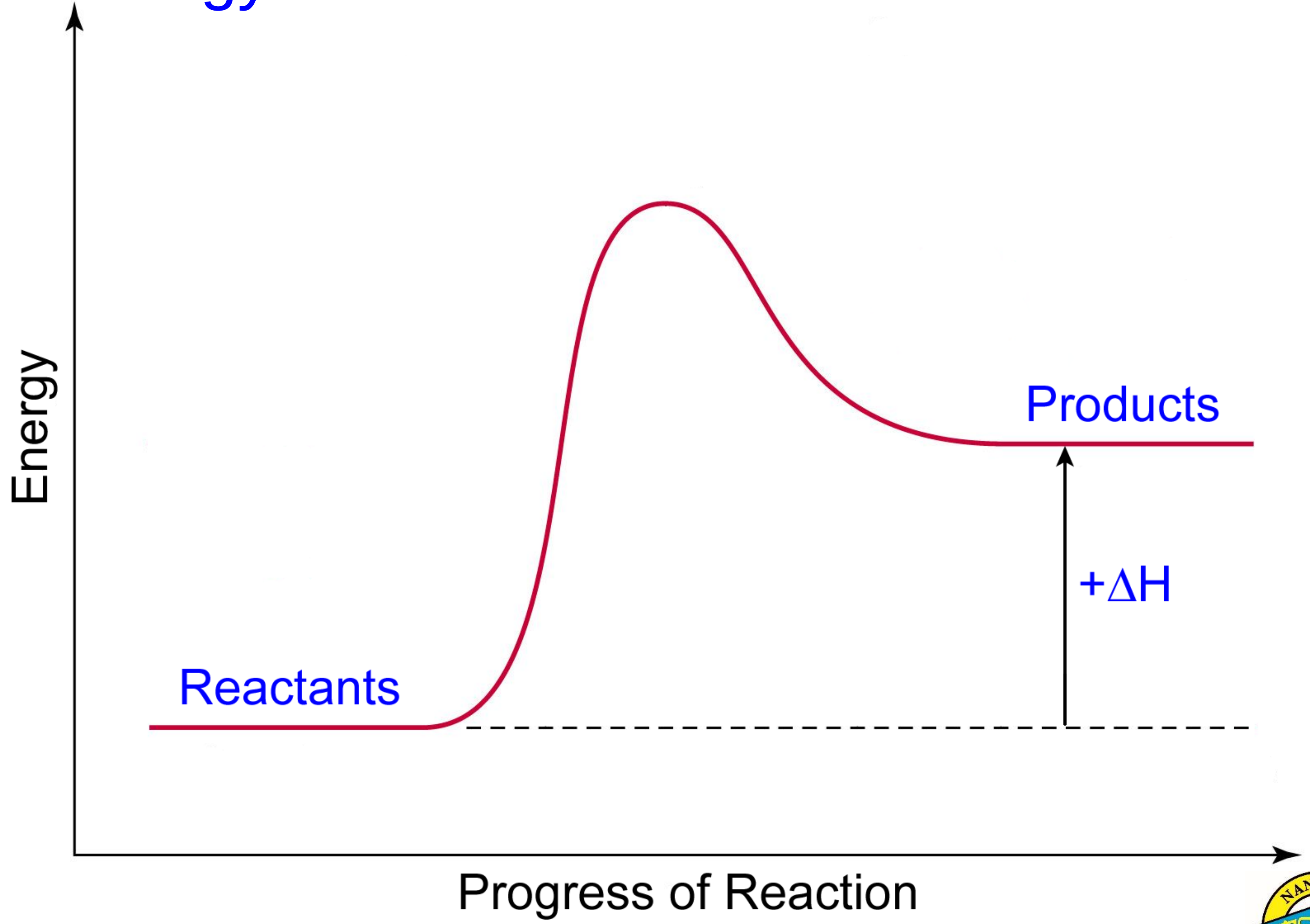
Reactants

Progress of Reaction

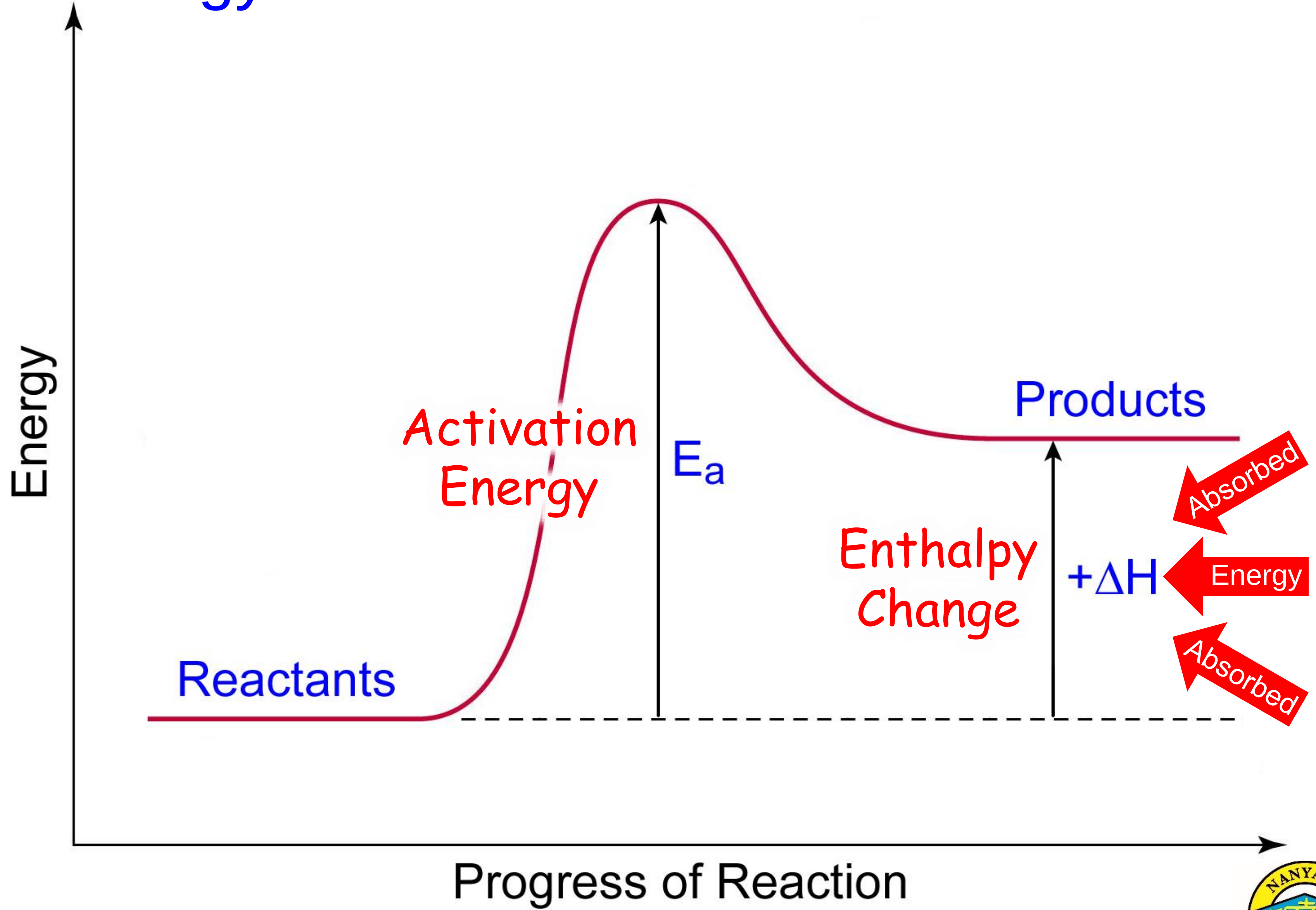
Energy Profile for an Endothermic Reaction



Energy Profile for an Endothermic Reaction



Energy Profile for an Endothermic Reaction



Energy Profile for an Endothermic Reaction

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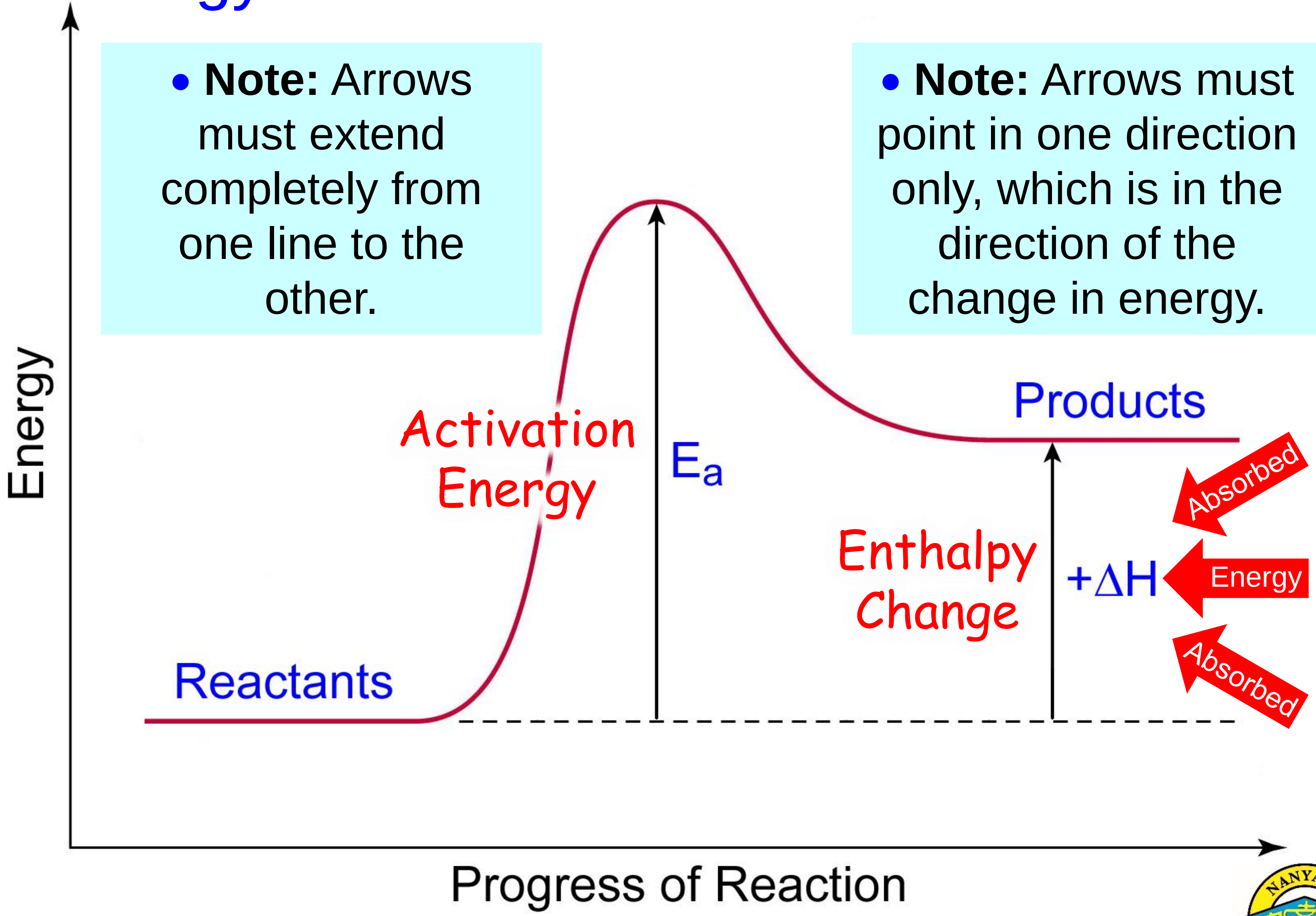
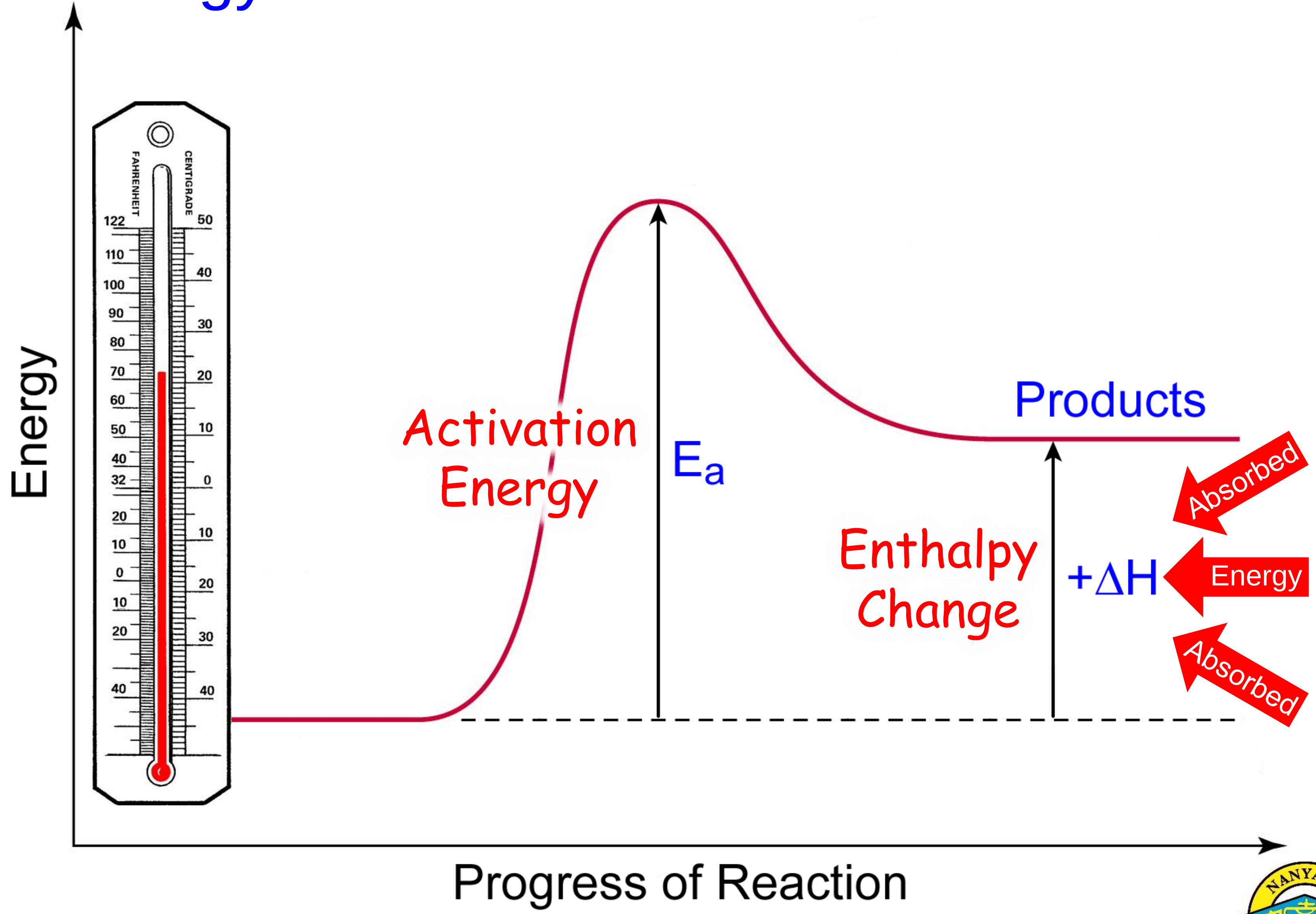
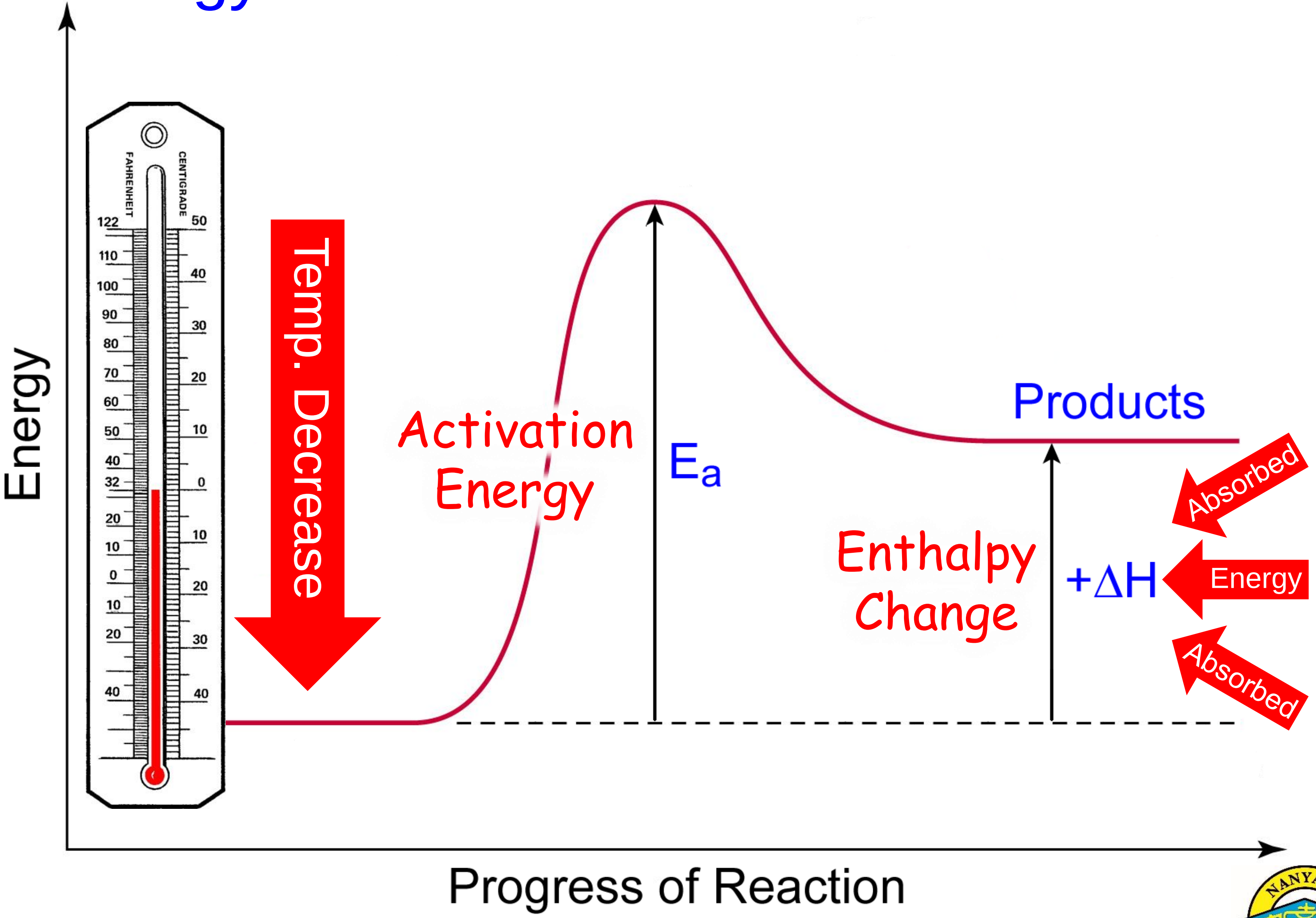


Energy Profile for an Endothermic Reaction



Energy Profile for an Endothermic Reaction



Energy

Temp. Decrease

Activation Energy

E_a

Products

Enthalpy Change

$+\Delta H$

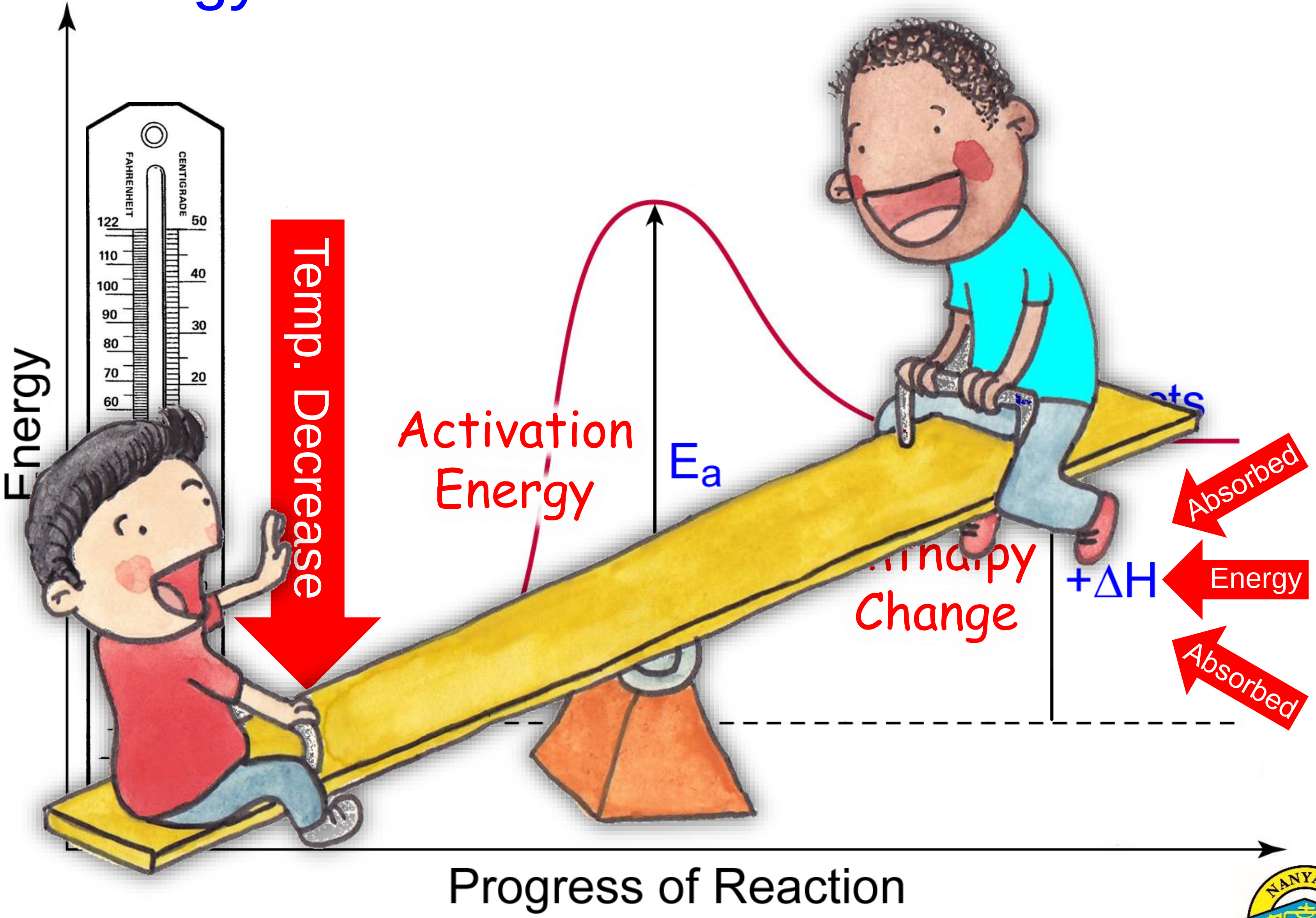
Absorbed

Energy

Absorbed

Progress of Reaction

Energy Profile for an Endothermic Reaction



Energy

Temp. Decrease

Activation Energy

E_a

ts

nthalpy Change

+ΔH

Absorbed

Energy

Absorbed

Progress of Reaction

Energy from Chemicals

Chemical reactions are accompanied by energy changes. When the energy content of a chemical system decreases during a reaction, the energy content of the products is ***greater / less*** than that of the reactants. The reaction is said to be ***exothermic / endothermic*** and a ***rise / fall*** in temperature is observed. The value of the enthalpy change of the reaction (ΔH) is ***positive / negative***.

When the temperature falls during a reaction, the reaction is said to be ***exothermic / endothermic****.* The energy content of the products is ***greater / less*** than that of the reactants *i.e.* the energy content of the system has ***increased / decreased*** during the reaction. ΔH for the reaction is ***positive / negative****.*

Energy from Chemicals

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When the temperature falls during a reaction, the reaction is said to be ~~***exothermic***~~ **/** ***endothermic.*** The energy content of the products is ***greater*** **/** ~~***less***~~ than that of the reactants *i.e.* the energy content of the system has ***increased*** **/** ~~***decreased***~~ during the reaction. ΔH for the reaction is ***positive*** **/** ~~***negative***~~.

Energy from Chemicals

What are some examples of *endothermic* reactions?

Endothermic Reactions

- Photosynthesis:

carbon dioxide + water → glucose + oxygen

$6CO_{2(g)} + 6H_2O_{(l)} \rightarrow C_6H_{12}O_{6(s)} + 6O_{2(g)}$

- Thermal Decomposition:

calcium carbonate → calcium oxide + carbon dioxide

$CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$

- Cracking Alkanes:

decane → octane + ethene

$C_{10}H_{22(l)} \rightarrow C_8H_{18(l)} + C_2H_{4(g)}$

Energy from Chemicals

Which type of chemical reaction is more common, *exothermic* or *endothermic*?

Exothermic and Endothermic Reactions

- Chemical reactions take place in order for chemicals to reduce their energy and become more stable.
- It is natural for chemicals to go from a state of high energy to a state of low energy, just as it is natural for a ball to roll down a hill.

Exothermic and Endothermic Reactions

- In an exothermic reaction, high energy reactants form low energy products, with energy given off to the surroundings.
- In an endothermic reaction, low energy reactants form high energy products, with energy absorbed from the surroundings.

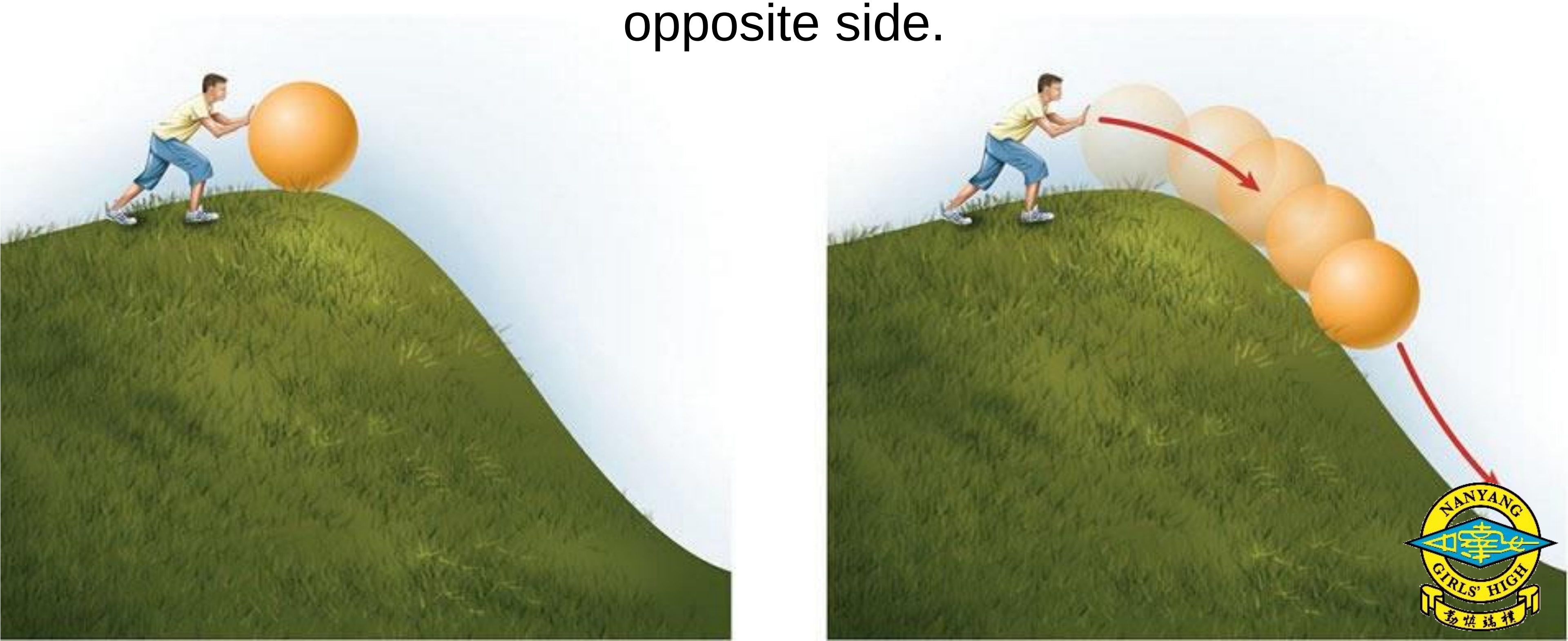
Exothermic and Endothermic Reactions

- In terms of energy changes, exothermic reactions are more favorable than endothermic reactions.
- Exothermic reactions are therefore more common than endothermic reactions.

Exothermic and Endothermic Reactions

- Question: How do these diagrams illustrate the concept of activation energy?

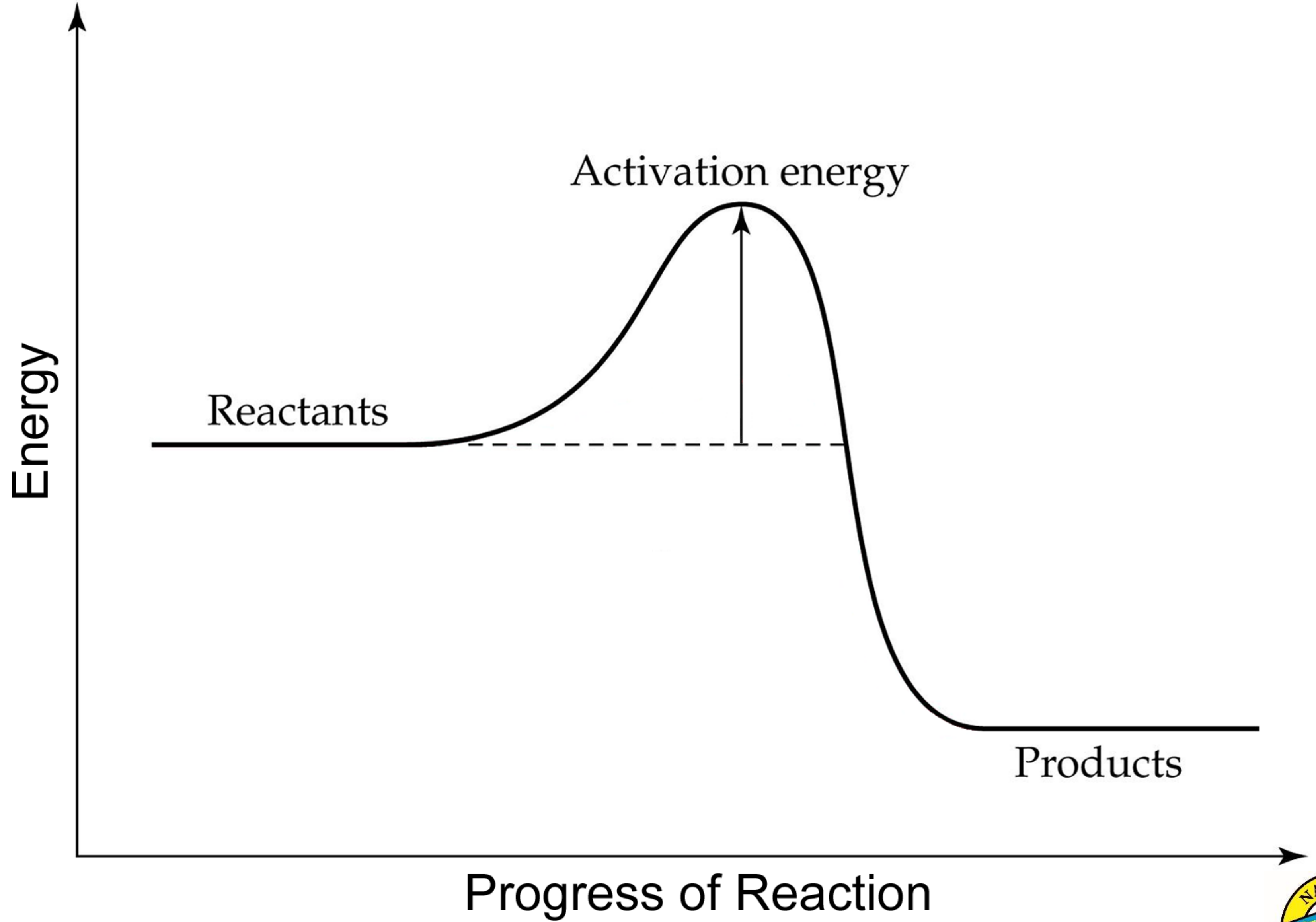
- Answer: The person must use a specific amount of energy to roll the ball from the left-hand-side to the top of the hill. Only then is the ball free to roll down the opposite side.



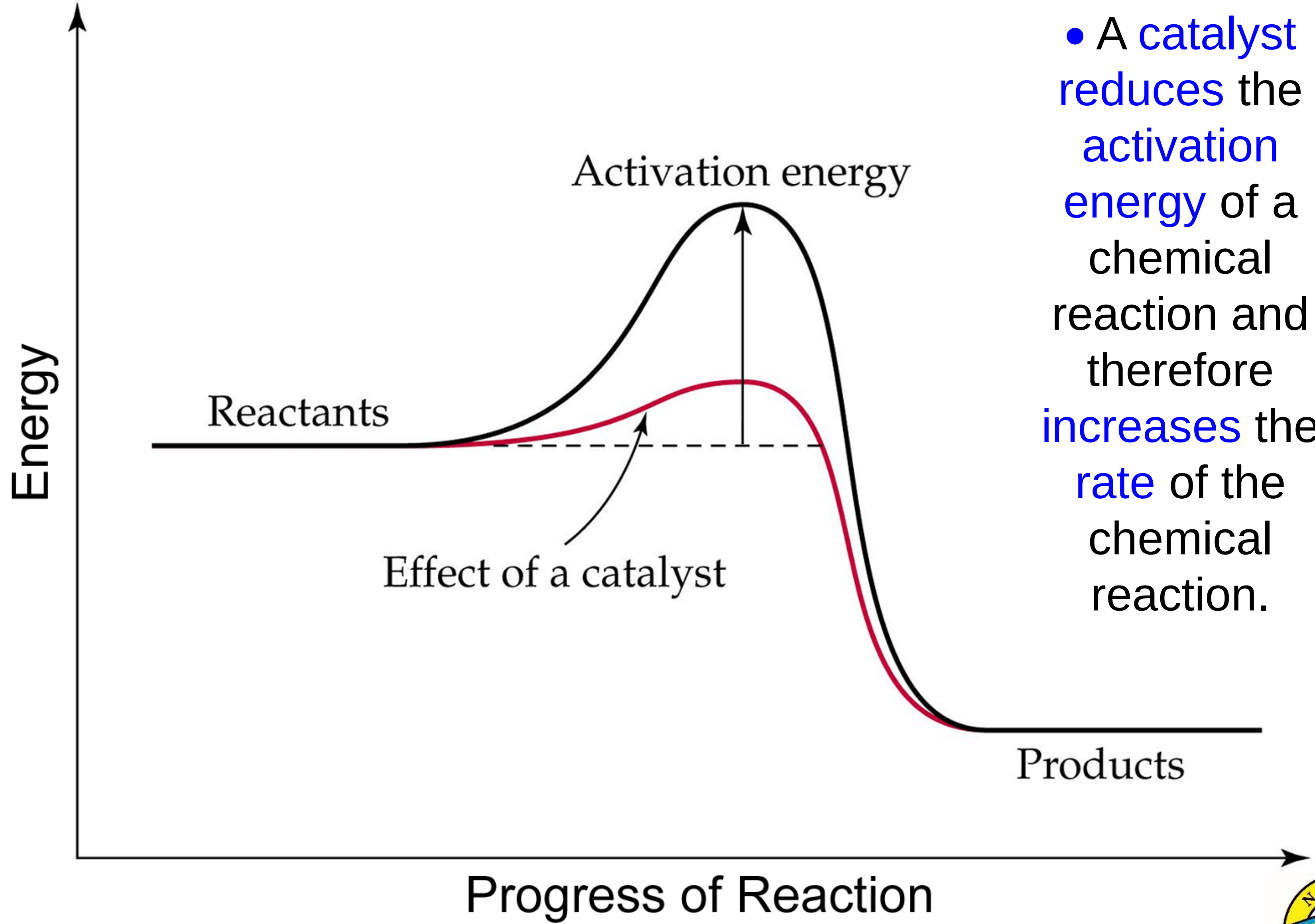
Energy from Chemicals

What effect does a *catalyst* have on the *activation energy* of a chemical reaction?

Effect of a Catalyst

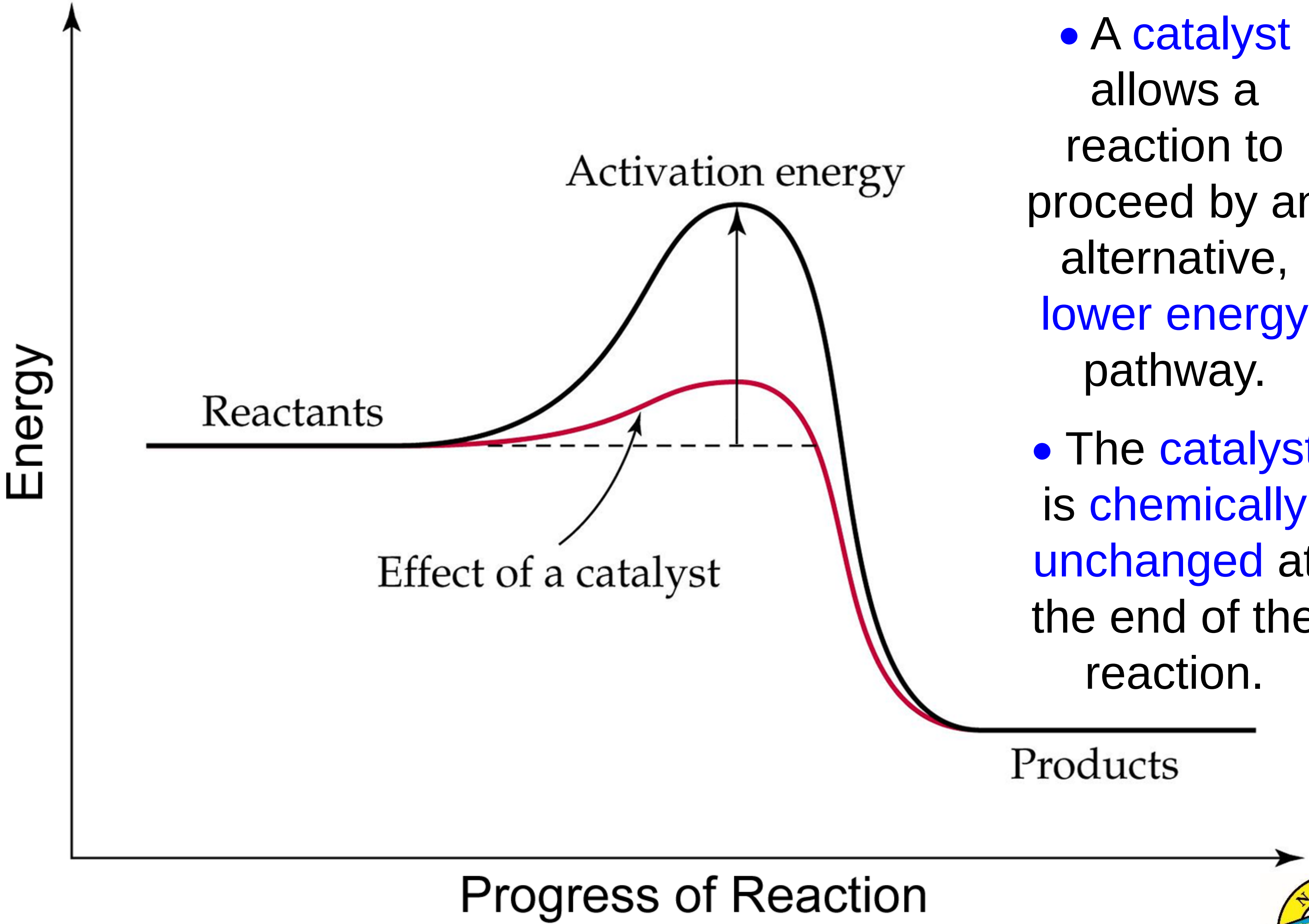


Effect of a Catalyst



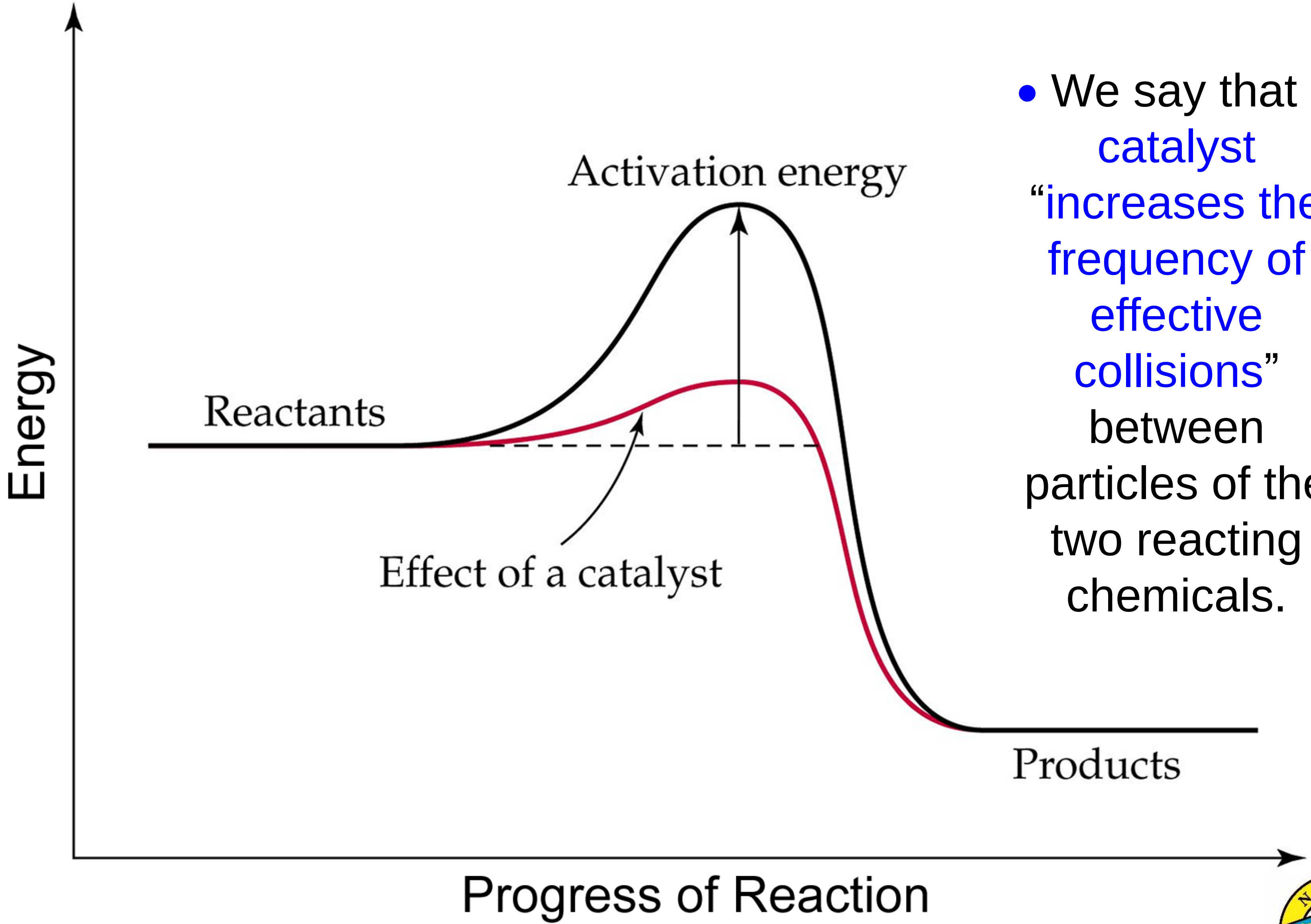
- A catalyst reduces the activation energy of a chemical reaction and therefore increases the rate of the chemical reaction.

Effect of a Catalyst



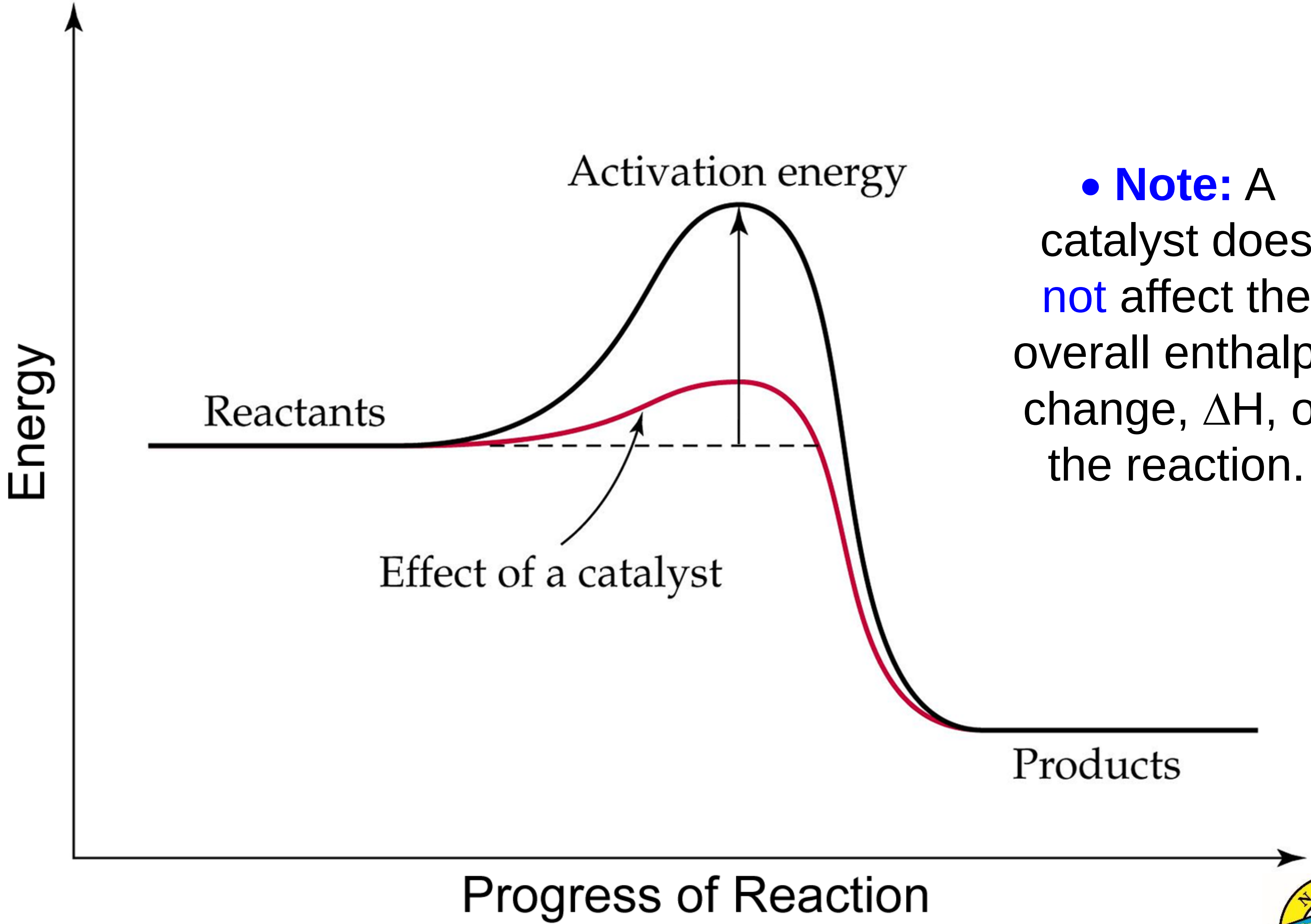
- A catalyst allows a reaction to proceed by an alternative, lower energy pathway.
- The catalyst is chemically unchanged at the end of the reaction.

Effect of a Catalyst



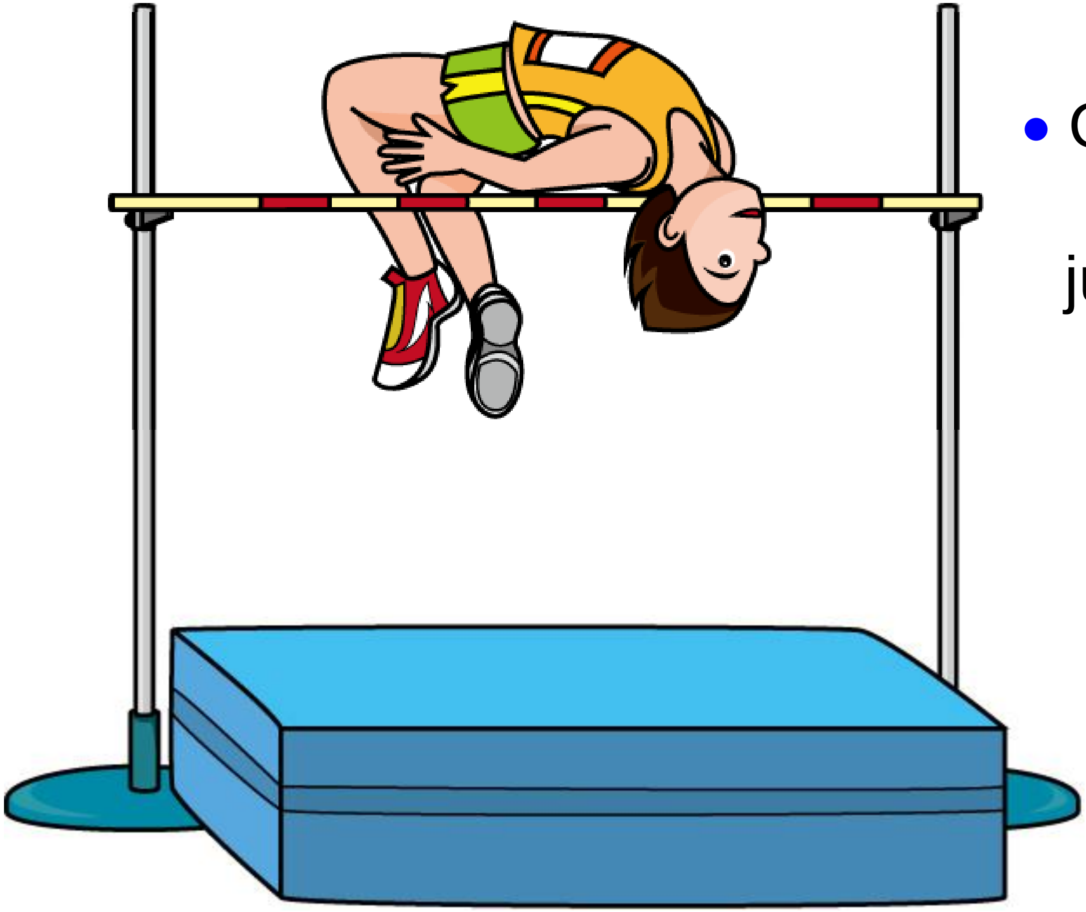
- We say that a catalyst "increases the frequency of effective collisions" between particles of the two reacting chemicals.

Effect of a Catalyst



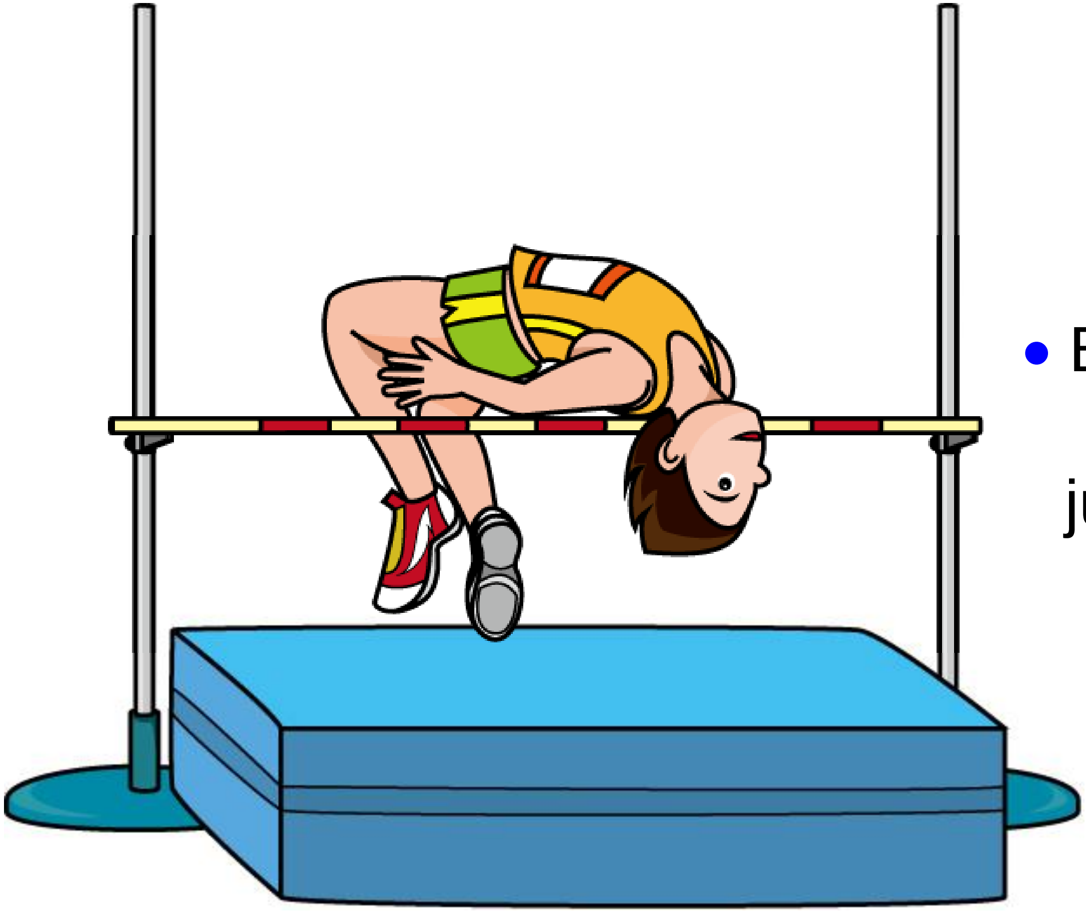
- **Note:** A catalyst does not affect the overall enthalpy change, ΔH, of the reaction.

Effect of a Catalyst



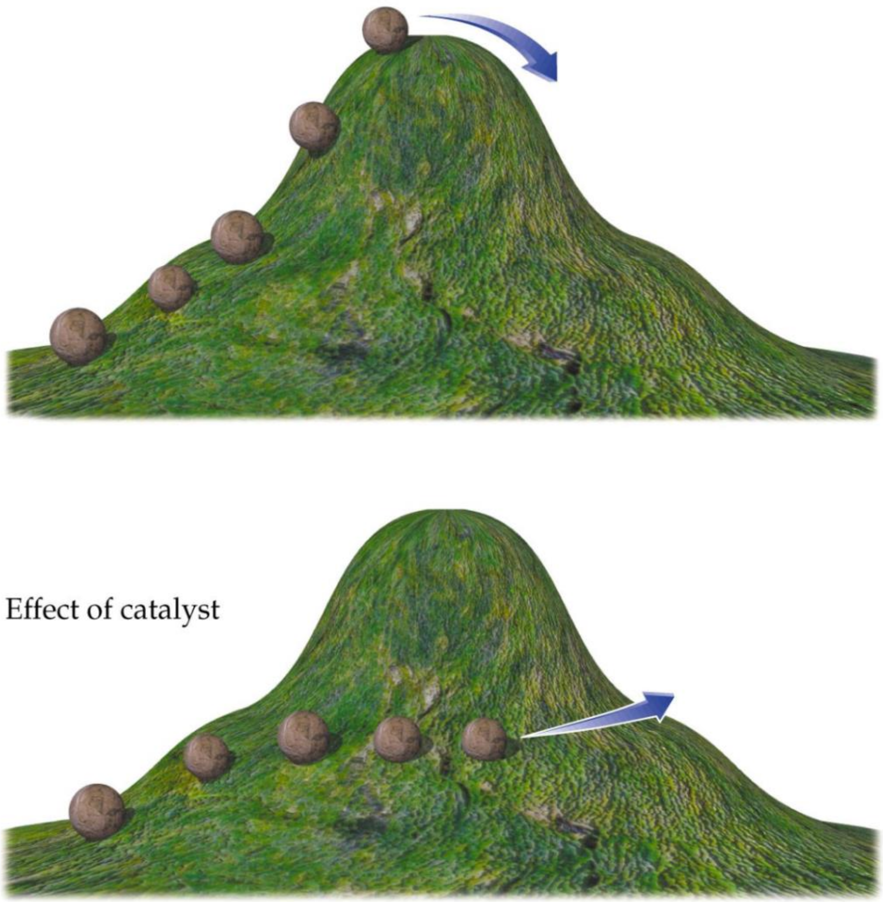
- Only 3 out of 30 students can jump this high.

Effect of a Catalyst



- But 24 out of 30 students can jump this high.

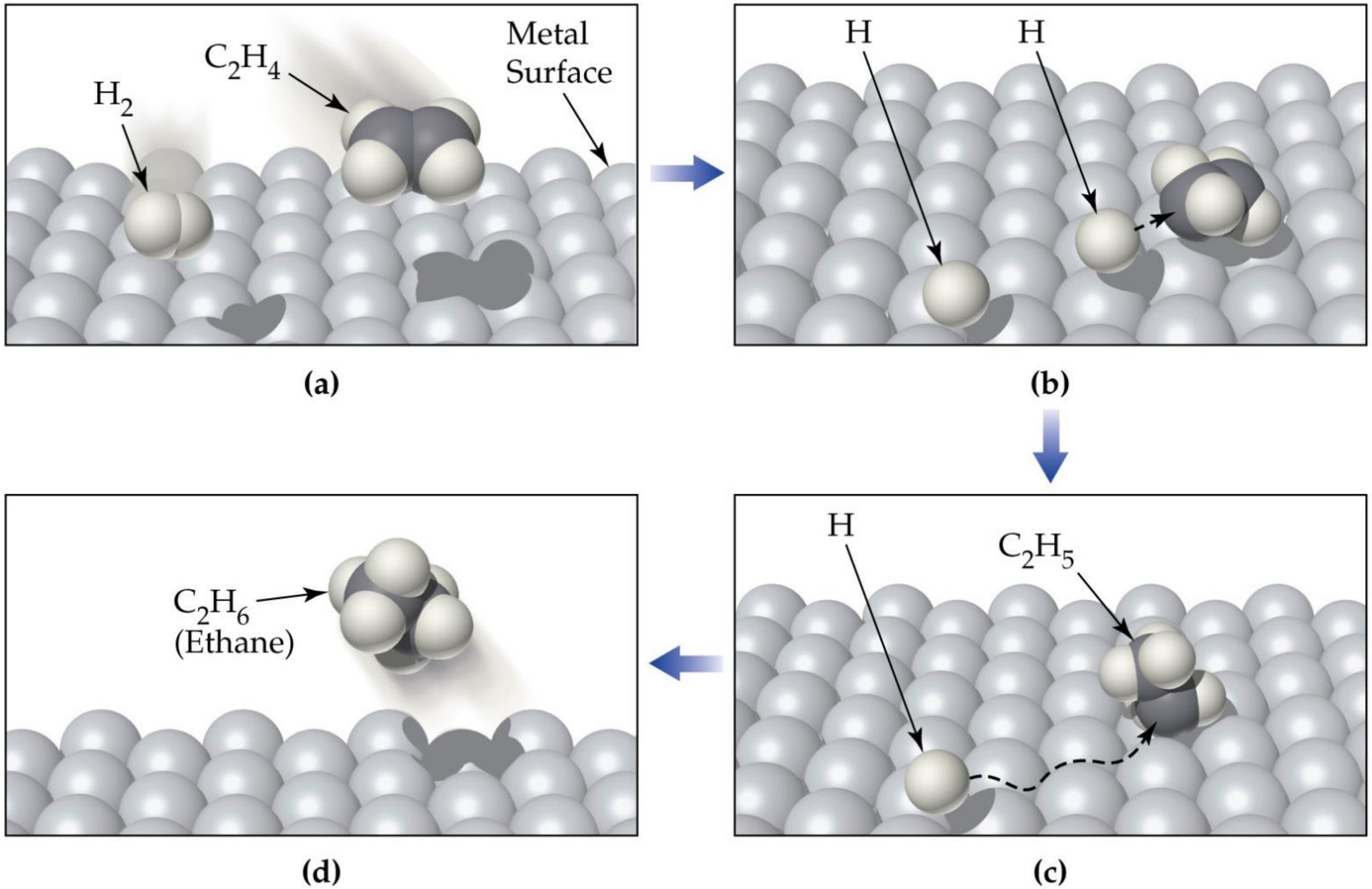
Effect of a Catalyst



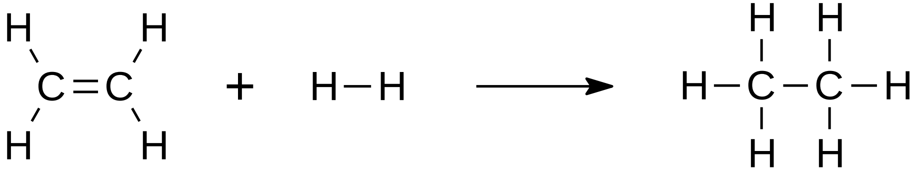
- Analogy for activation energy. Less energy is required to push the boulders around the side of the hill compared to the energy that is required to push the boulders over the hill.

- Using the same amount of energy, more boulders can be pushed around the side of the hill compared to the number that can be pushed over the hill.

Effect of a Catalyst



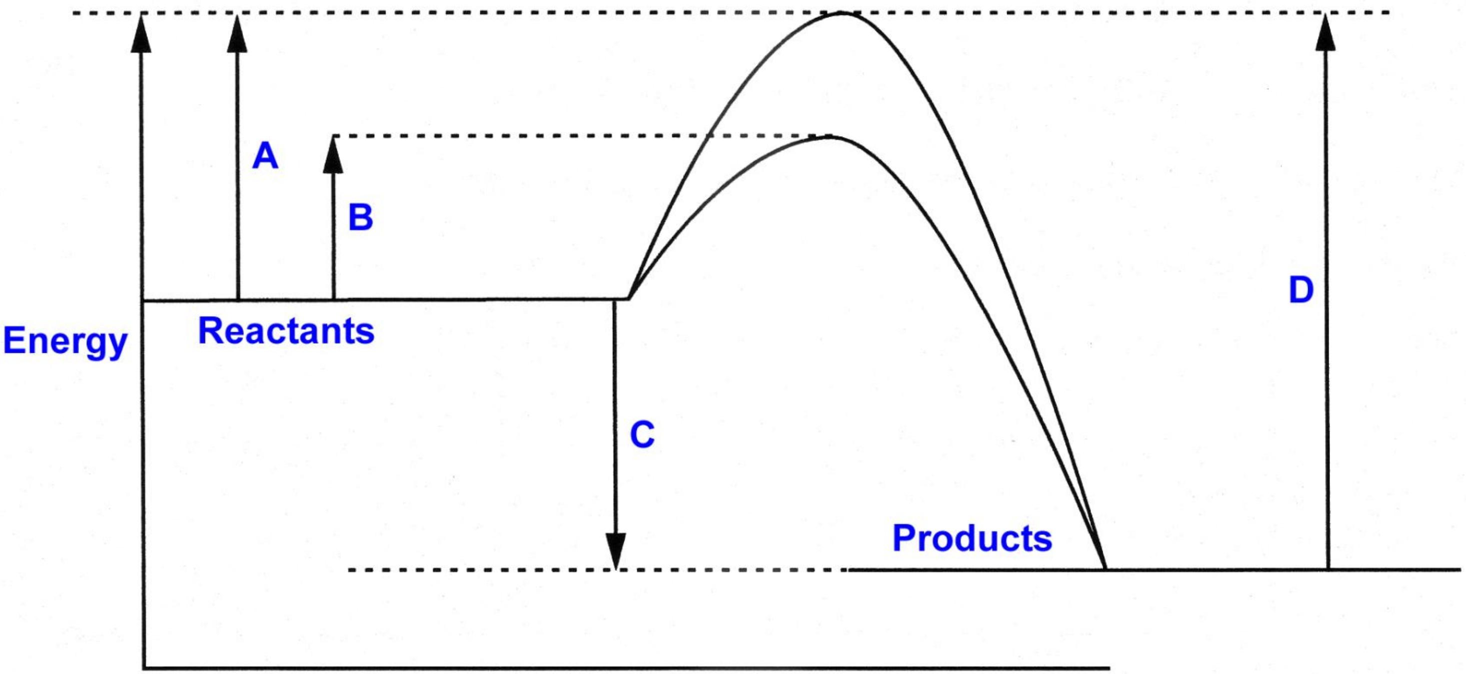
(a) **(b)** **(c)** **(d)**



Ethene Hydrogen Ethane

Assessment for Learning

- Study the energy change diagram shown below.
- Identify energy changes **A**, **B**, **C** and **D**.



Assessment for Learning

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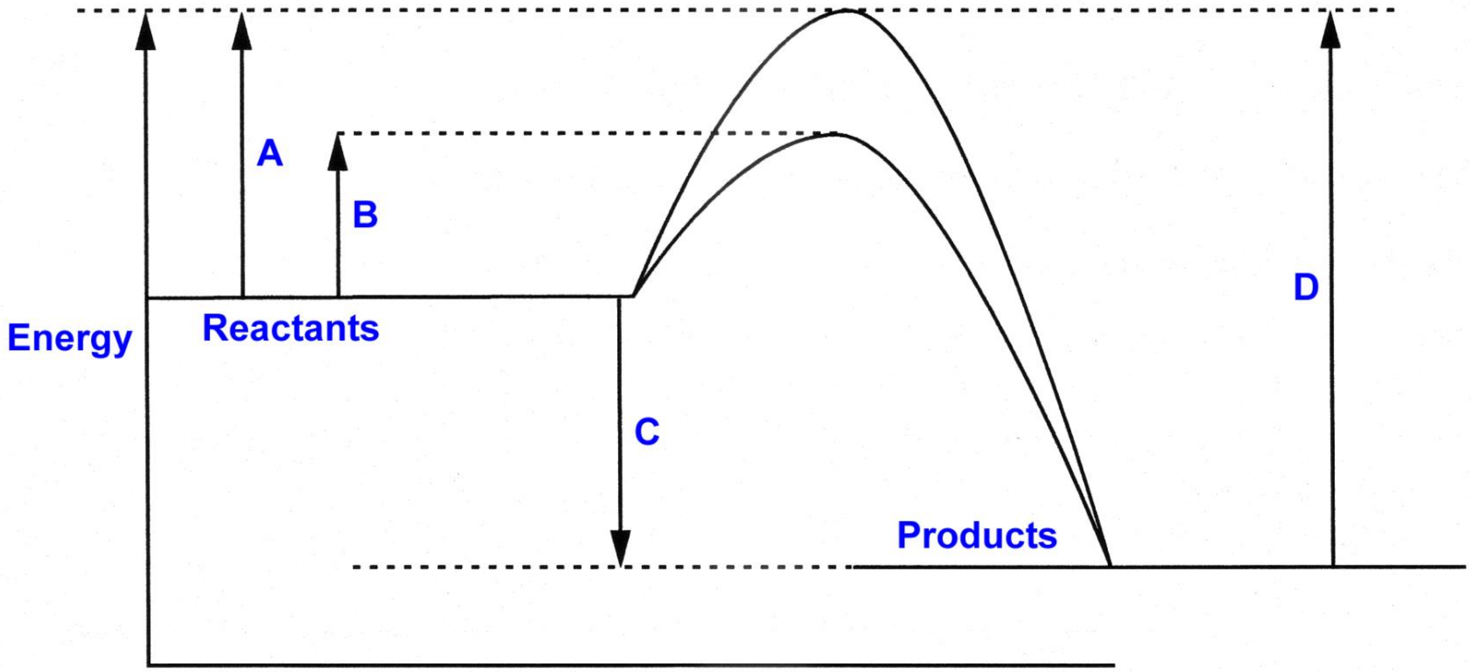


A Activation energy of the forward reaction.

B Activation energy of the forward reaction with a catalyst.

C Enthalpy change (ΔH) of the forward reaction.

D Activation energy of the backward reaction.

Energy from Chemicals

How can I *calculate* the overall energy change of a reaction?

Energy from Chemicals

- Imagine that you were given \$50 (+50).
- You spend \$30 (–30).
- You will be in *credit* by \$20 (+50) + (–30) = +20.
- Overall, you have "*gained money from the surroundings*".

Energy from Chemicals

- Imagine that you were given \$50 (+50).
- You spend \$80 (–80).
- You will in *debt* by \$30 (+50) + (–80) = –30.
- Overall, you have "*lost money to the surroundings*".

Energy Change Calculations

Reactions can be considered to take place in two stages:

1) Bonds between atoms in the reactants are *broken*.

2) New bonds between atoms are *formed* to produce the products.

Energy Change Calculations

Important Things to Note…

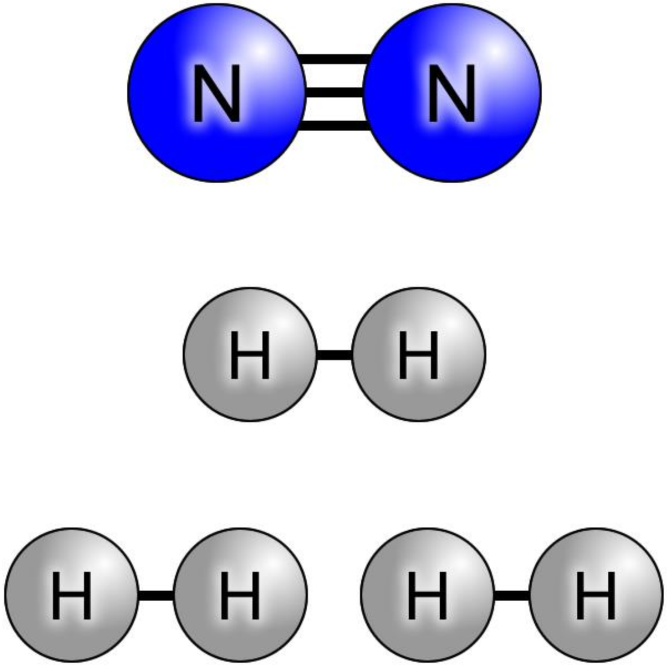
- Bond breaking *absorbs energy* and is therefore endothermic. Numerical values of ΔH for bond breaking are positive.

- Bond formation *releases energy* and is therefore exothermic. Numerical values of ΔH for bond formation are negative.

Energy Change Calculations

nitrogen + hydrogen $\rightleftharpoons$ ammonia

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$



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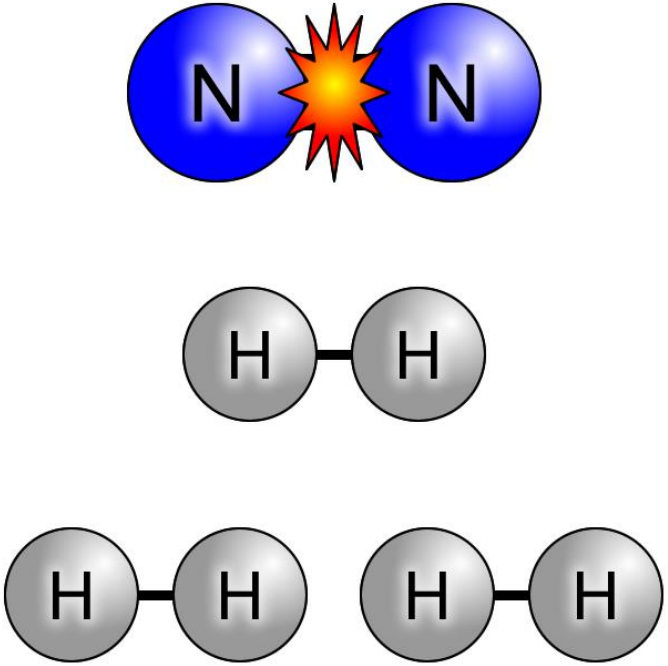

1) Bond breaking *absorbs* energy.
ΔH is positive.

Energy Change Calculations

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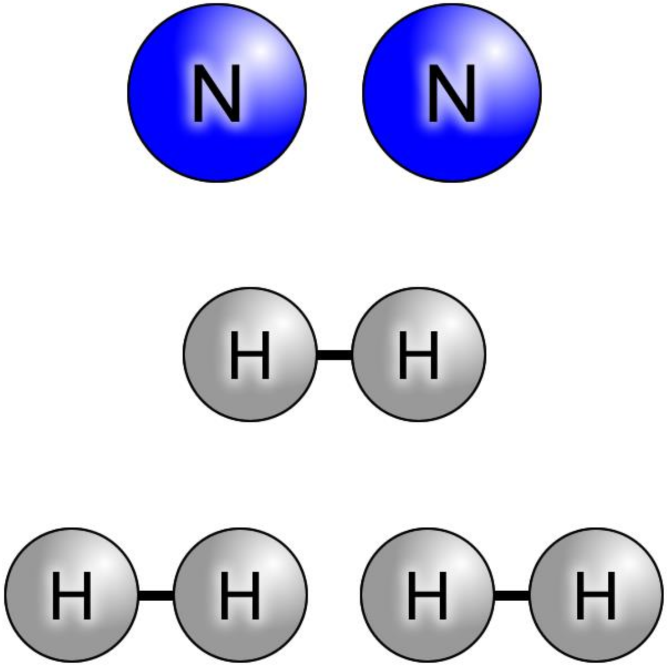

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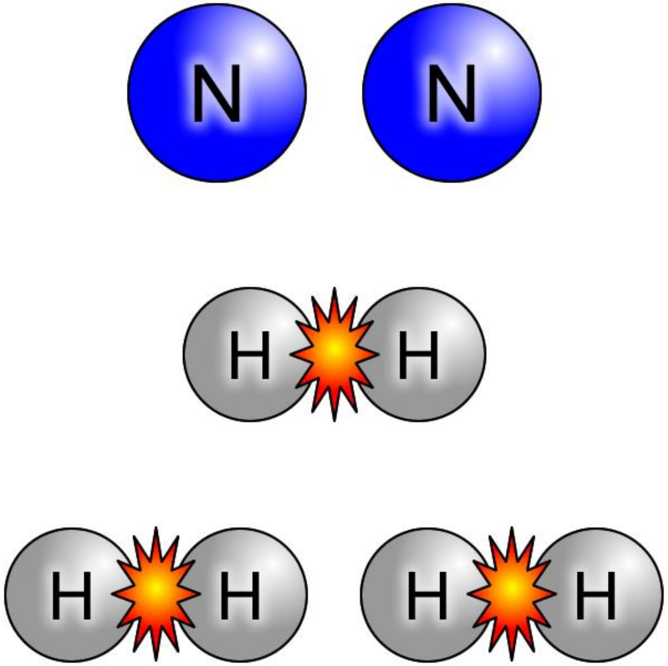

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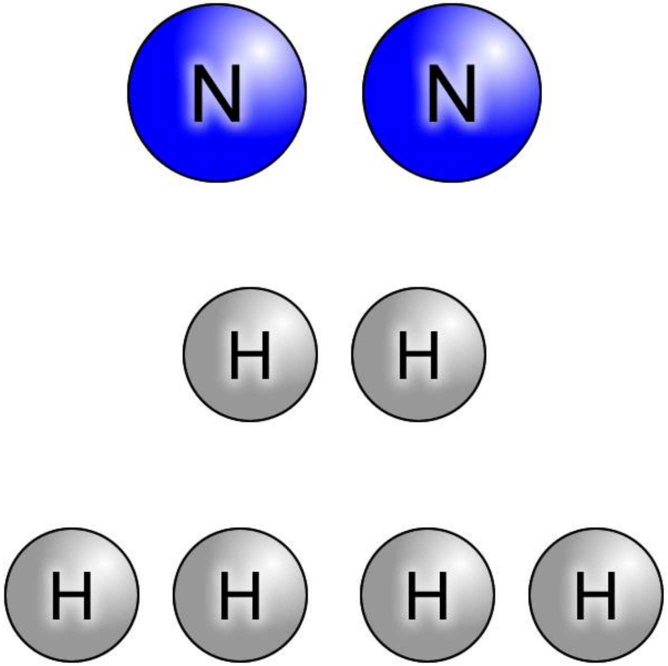

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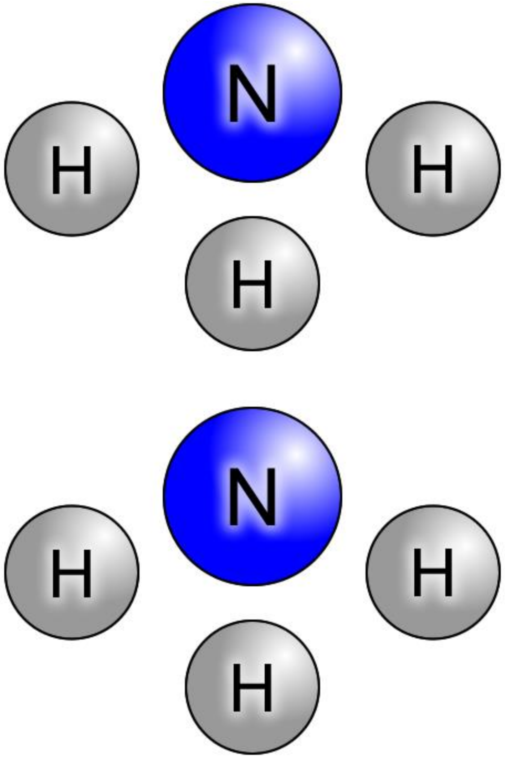

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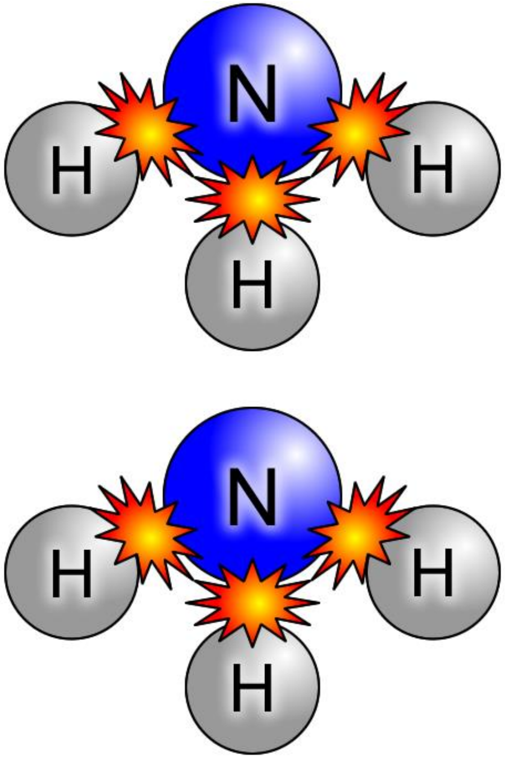
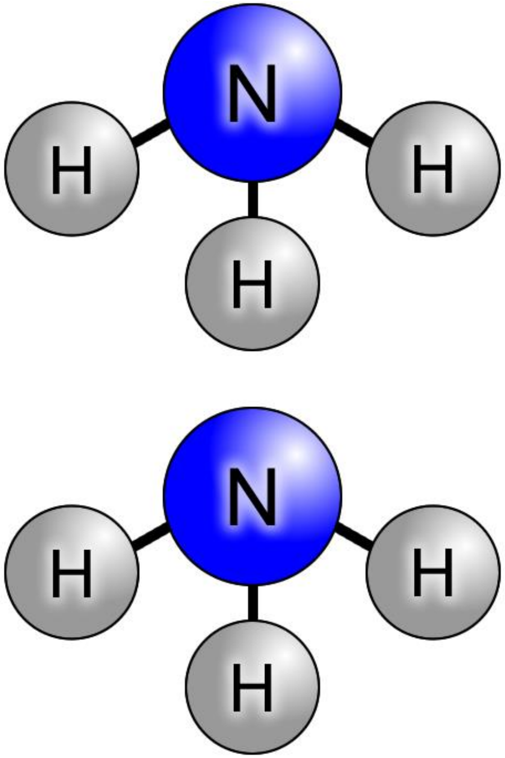


2) Bond formation
releases energy.
ΔH is negative.

Energy Change Calculations

nitrogen + hydrogen ⇌ ammonia

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$



Energy Change Calculations

Important Things to Note…

Overall enthalpy change of a chemical reaction (ΔH) = Total amount of energy absorbed to break existing chemical bonds (+ΔH) + Total amount of energy released when new chemical bonds are formed (−ΔH)

Energy Change Calculations

If…

Total amount of energy absorbed to break existing chemical bonds $(+\Delta H)$

>

Total amount of energy released when new chemical bonds are formed $(-\Delta H)$

Overall, **energy is absorbed** and the reaction will be **endothermic**, *i.e.* ΔH will be positive.

Energy Change Calculations

- *Energy absorbed* to *break* bonds in the reactants is *greater* than the *energy released* when new bonds are *formed* in the products.

- Overall, the reaction *absorbs energy* from the surroundings. The reaction is therefore *endothermic*.

Energy Change Calculations

If…

Total amount of energy absorbed to break existing chemical bonds ($+\Delta H$)

$<$

Total amount of energy released when new chemical bonds are formed ($-\Delta H$)

Overall, **energy is released** and the reaction will be **exothermic**, *i.e.* ΔH will be negative.

Energy Change Calculations

- *Energy absorbed* to *break* bonds in the reactants is *less* than the *energy released* when new bonds are *formed* in the products.

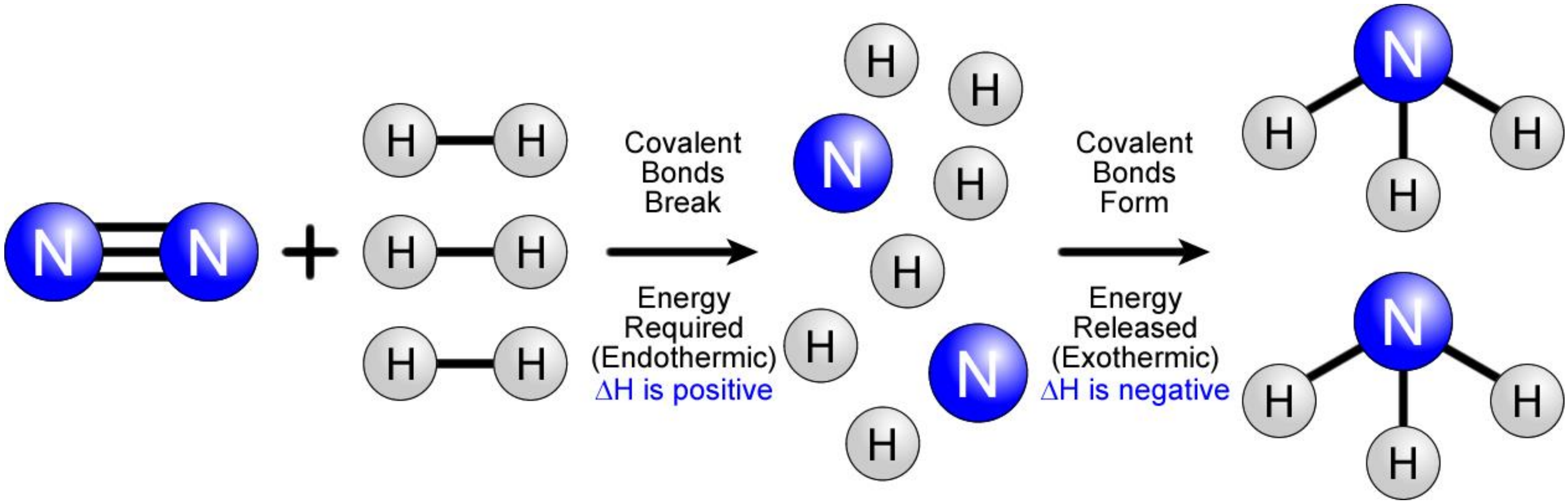
- Overall, the reaction *releases energy* into the surroundings. The reaction is therefore *exothermic*.

Energy Change Calculations – Example #1:

The Reaction Between Nitrogen and Hydrogen to form Ammonia

$$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$$

Overview of the Reaction:



Energy Change Calculations – Example #1:

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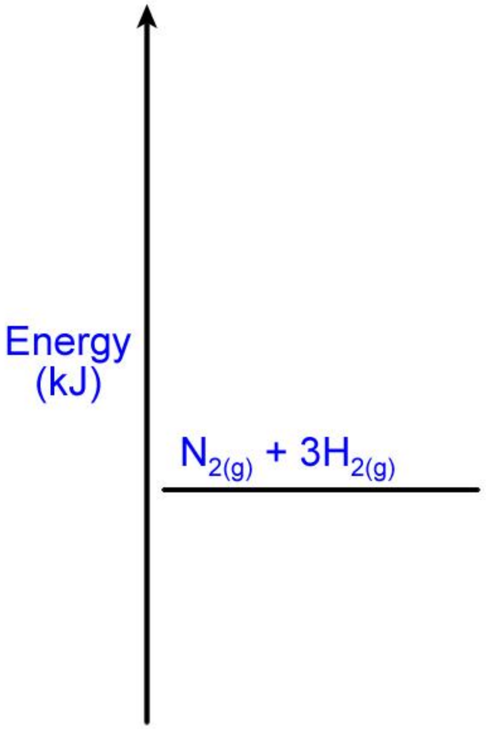
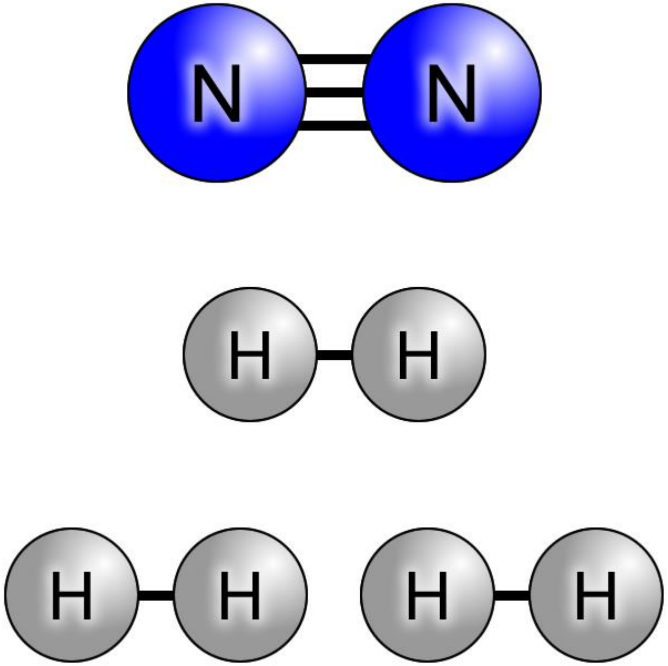
N≡N = 944 kJ/mol

H–H = 436 kJ/mol

N–H = 388 kJ/mol





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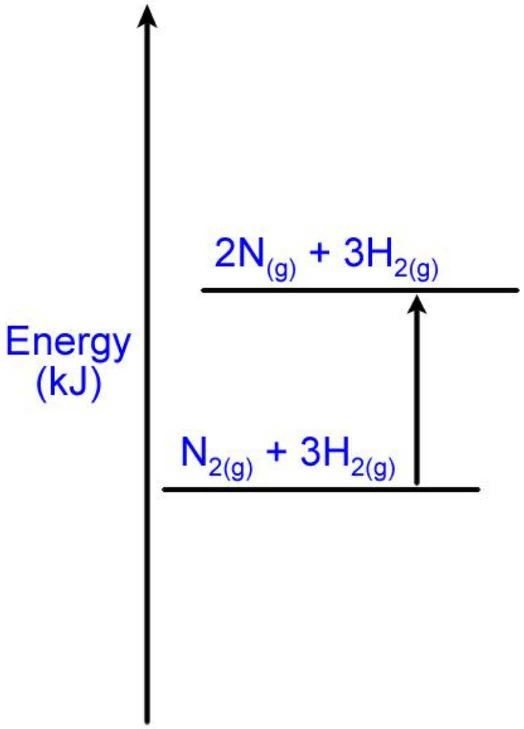

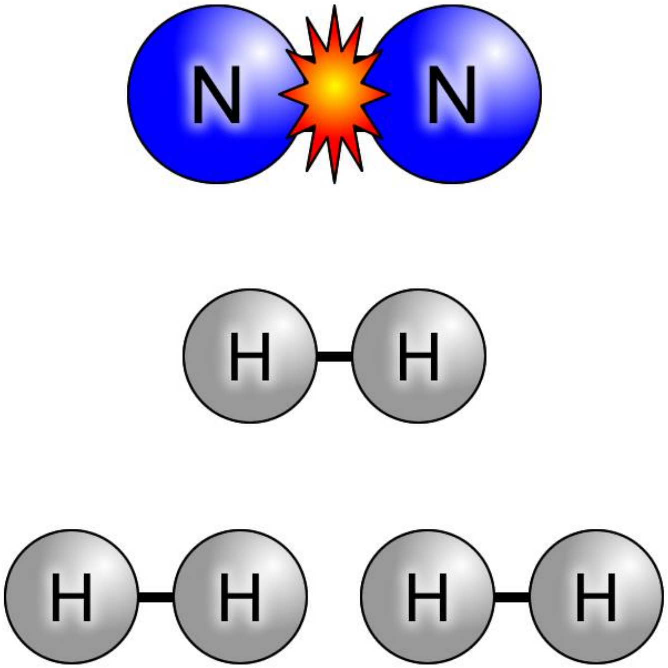

1 × N≡N bond is broken forming 2 × N atoms.

Energy Change Calculations – Example #1:

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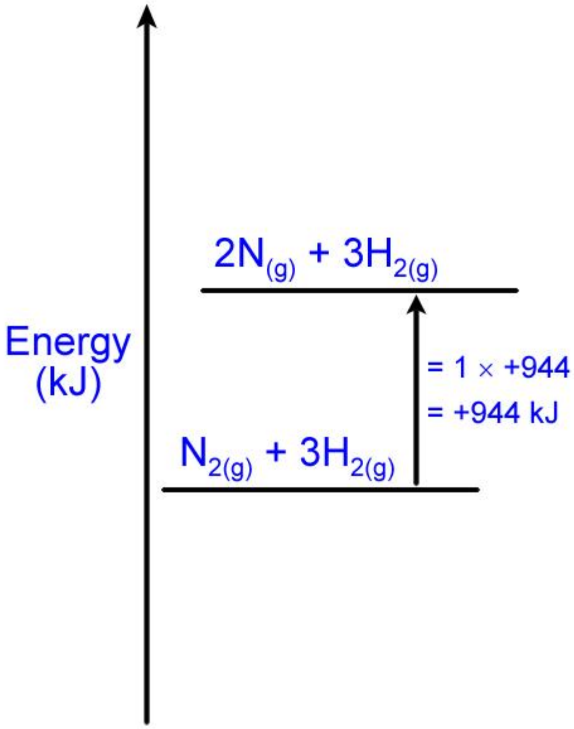

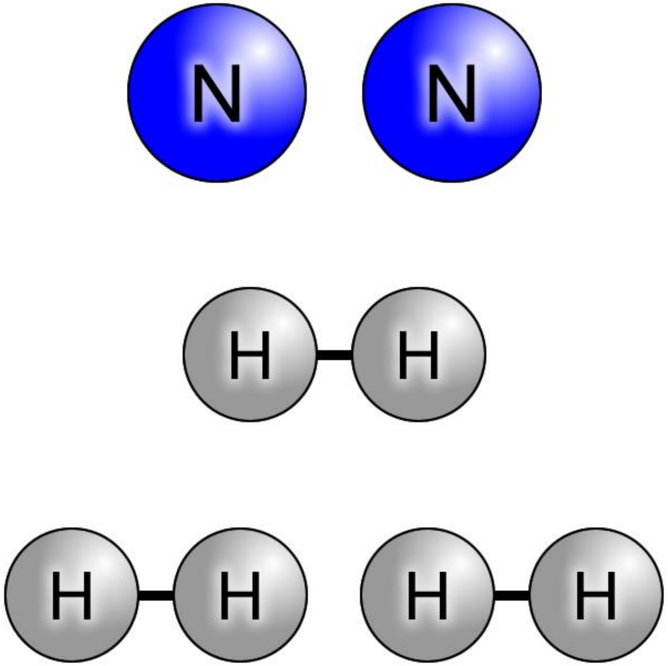

Bond breaking is *endothermic*: ΔH for this change is *positive*.

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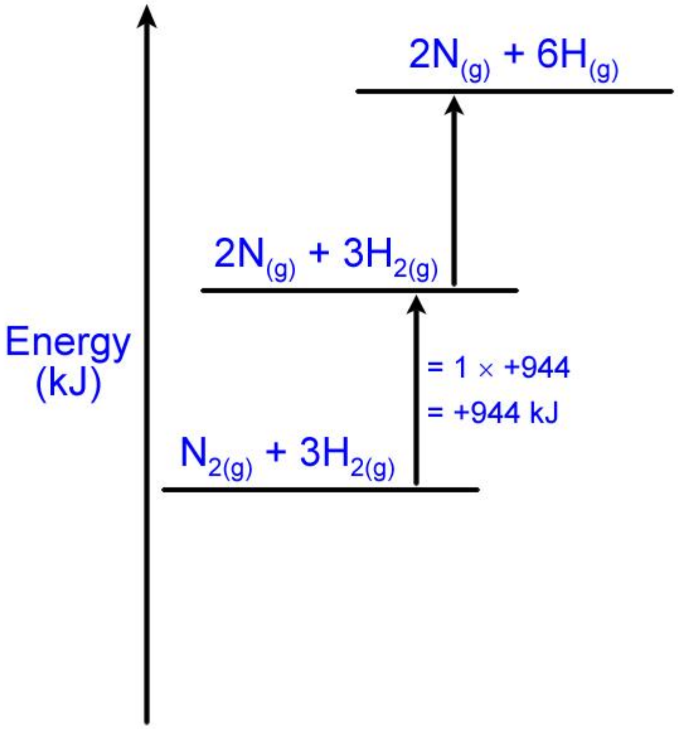

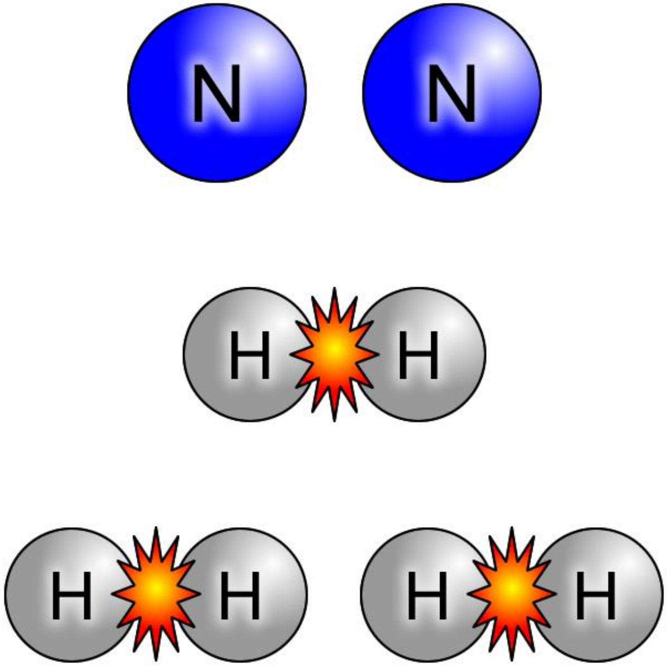

3 × H–H bonds are broken forming 6 × H atoms.

Energy Change Calculations – Example #1:

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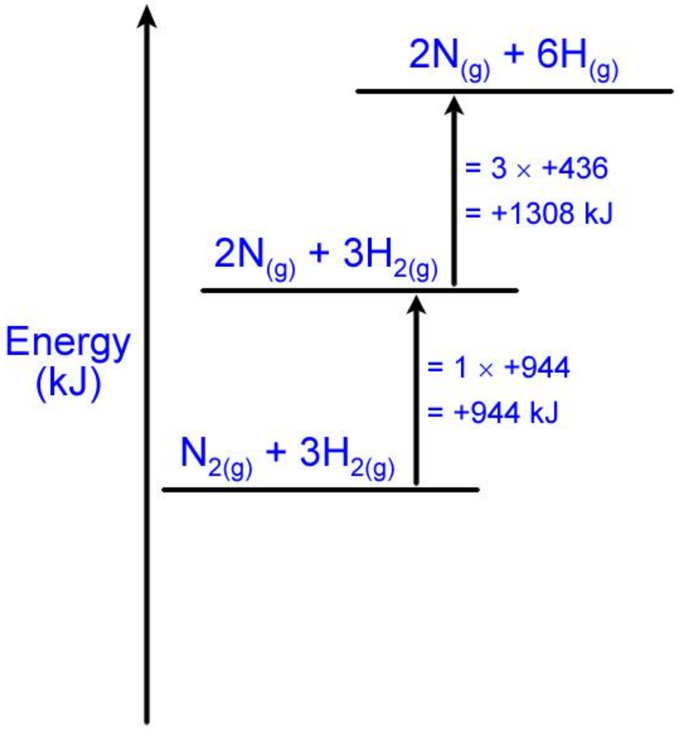

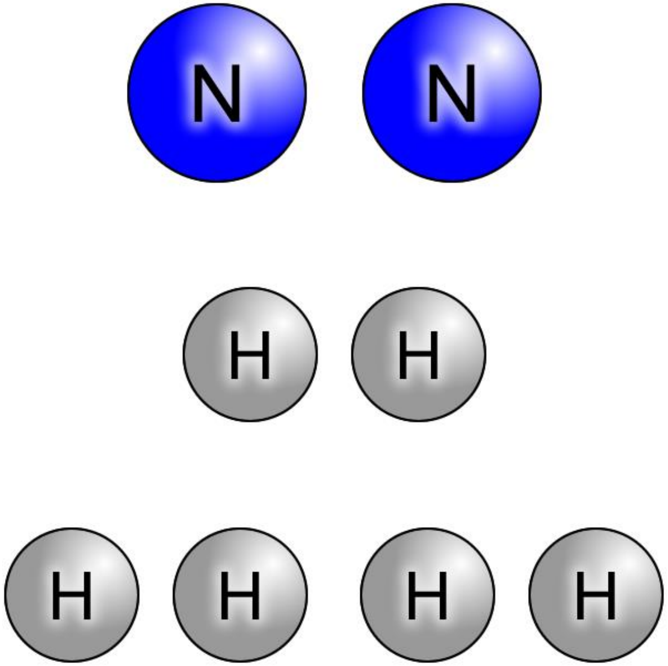

Bond breaking is *endothermic*: ΔH for this change is *positive*.

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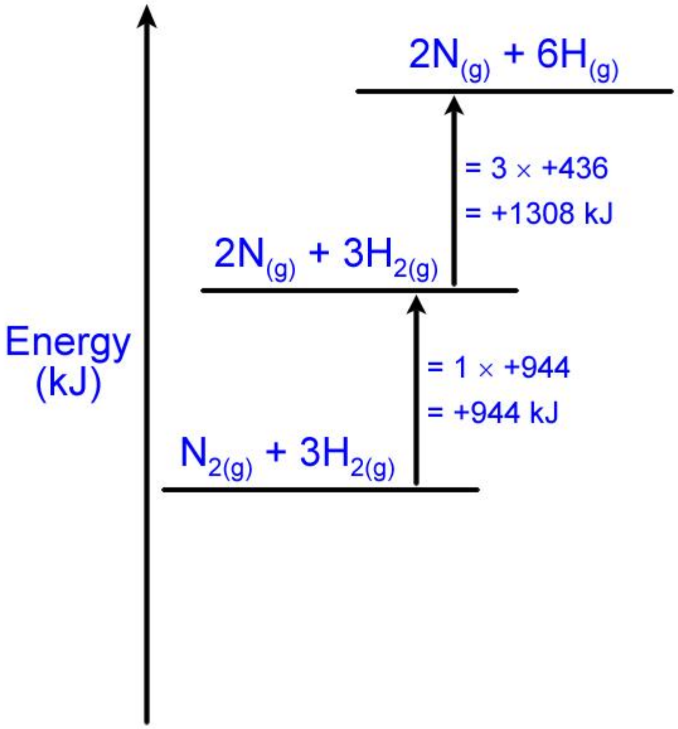

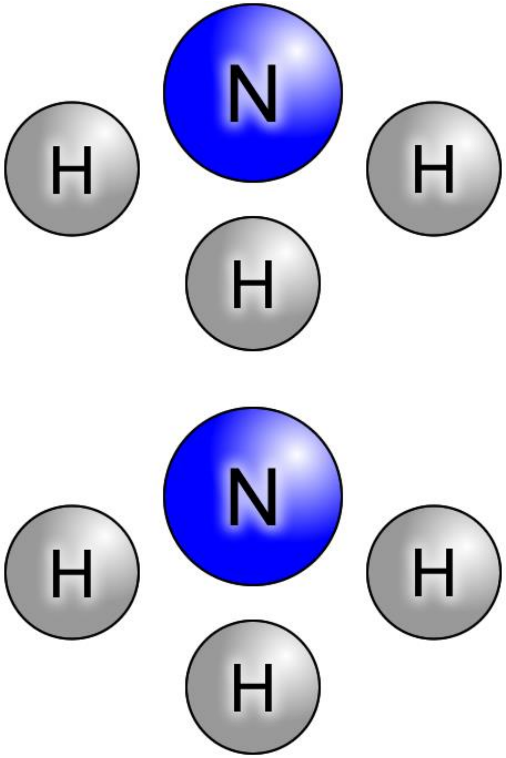

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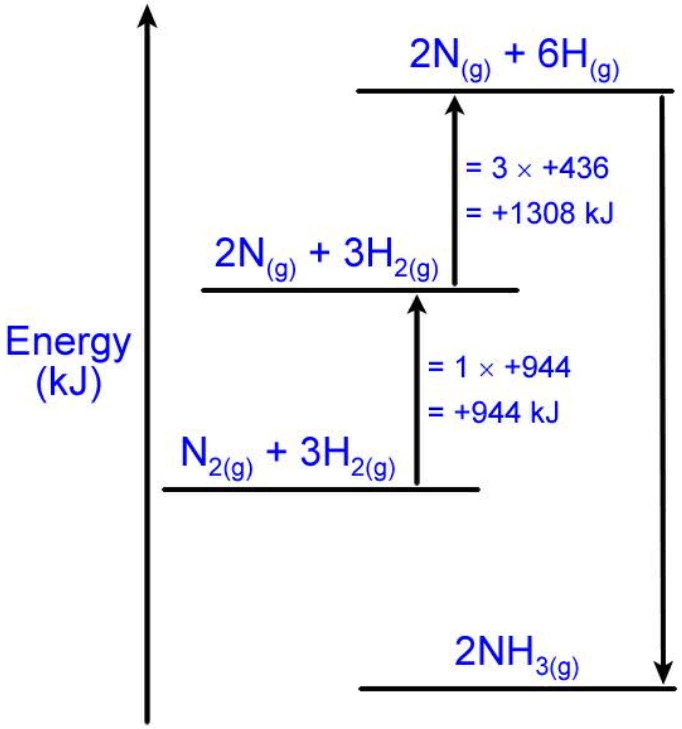

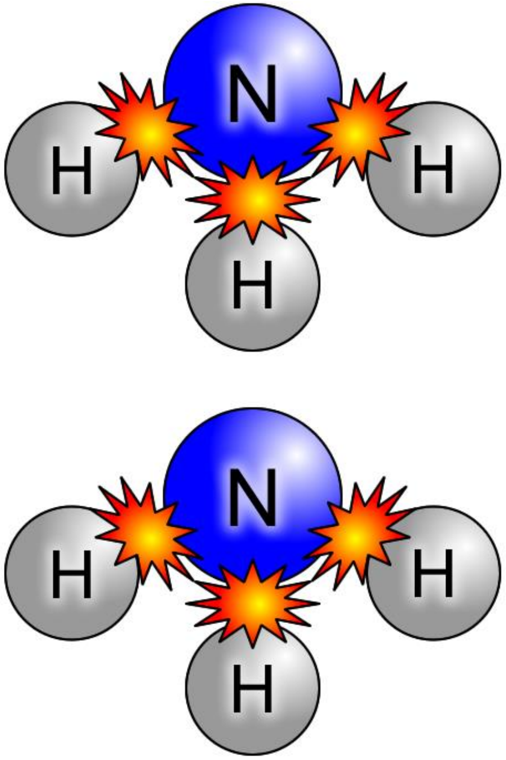

6 × N–H bonds are formed creating 2 × NH_3 molecules.

Energy Change Calculations – Example #1:

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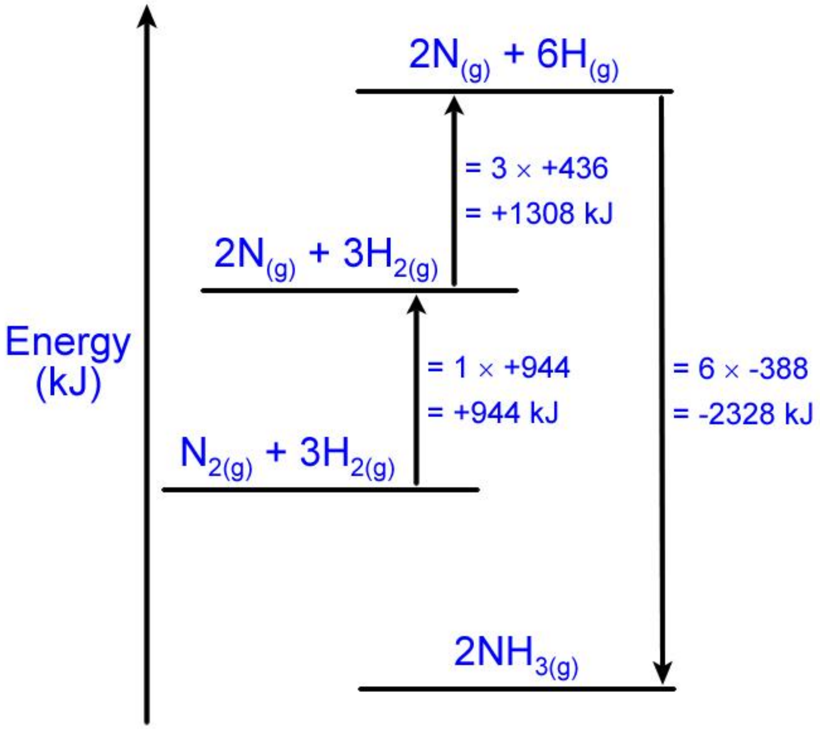

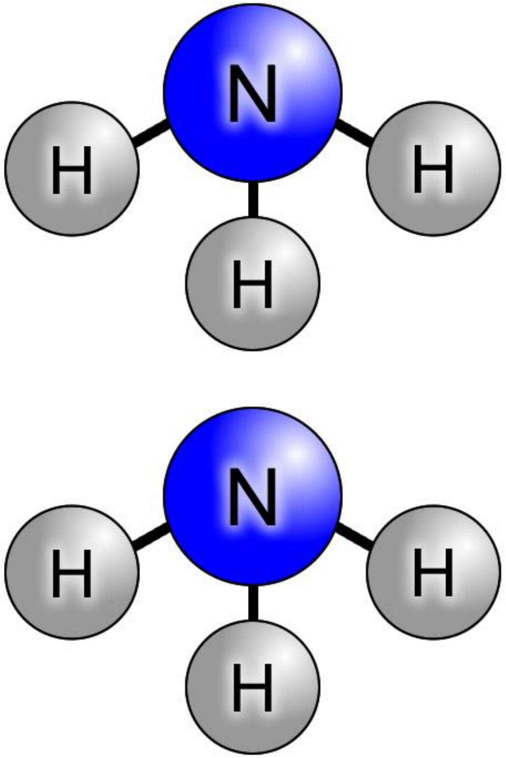

Bond formation is *exothermic*: ΔH for this change is *negative*.

Energy Change Calculations – Example #1:
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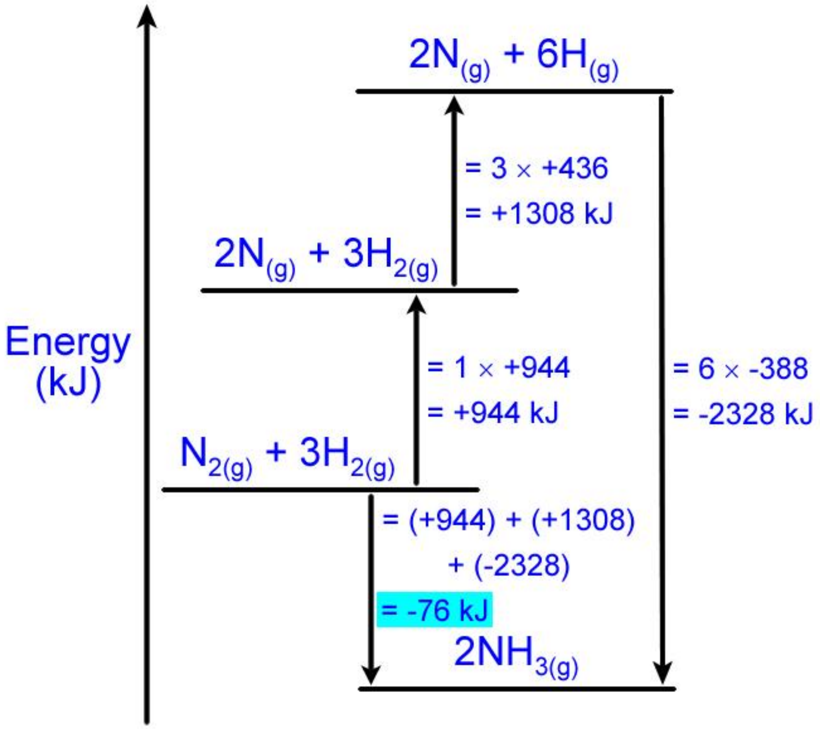

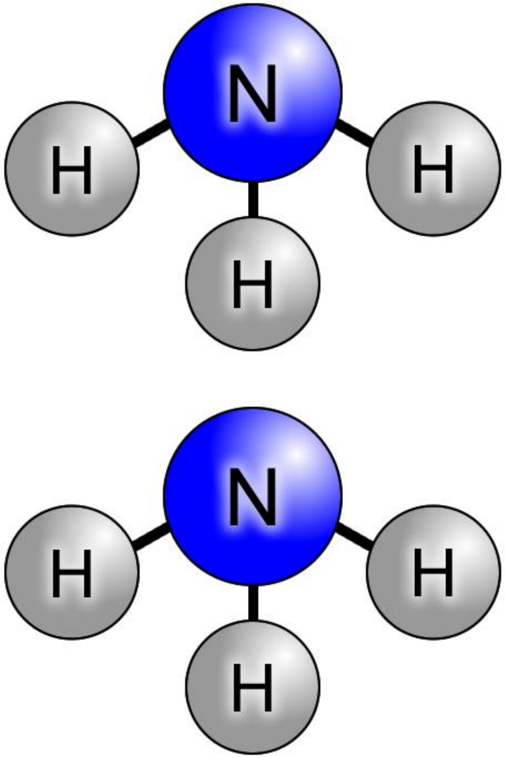

$$(+944) + (+1308) + (-2328) = -76.0\ kJ$$

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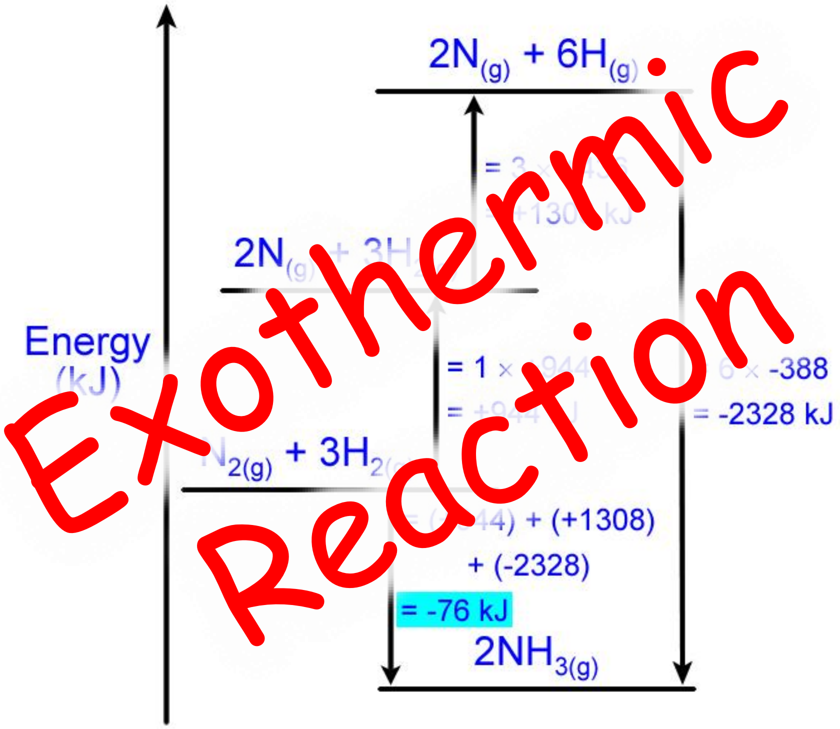
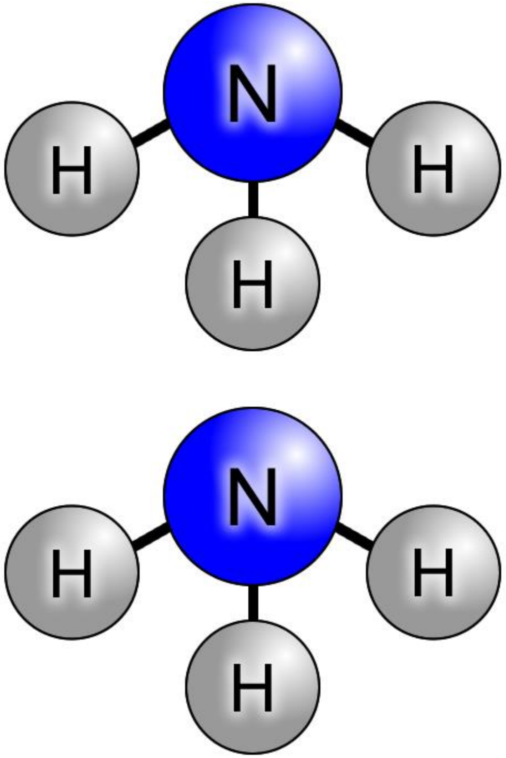


Exothermic Reaction



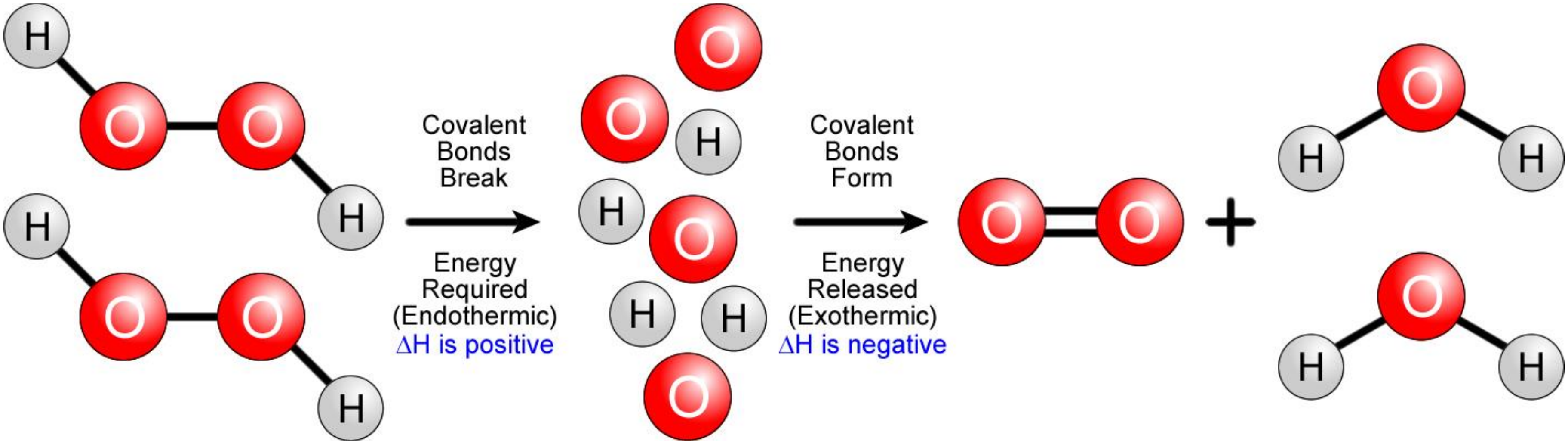
$$(+944) + (+1308) + (-2328) = \boxed{-76.0\ \mathrm{kJ}}$$

Energy Change Calculations – Example #2:

The Decomposition of Hydrogen Peroxide to form Oxygen and Water

$2H_2O_{2(g)} \rightarrow O_{2(g)} + 2H_2O_{(g)}$

Overview of the Reaction:



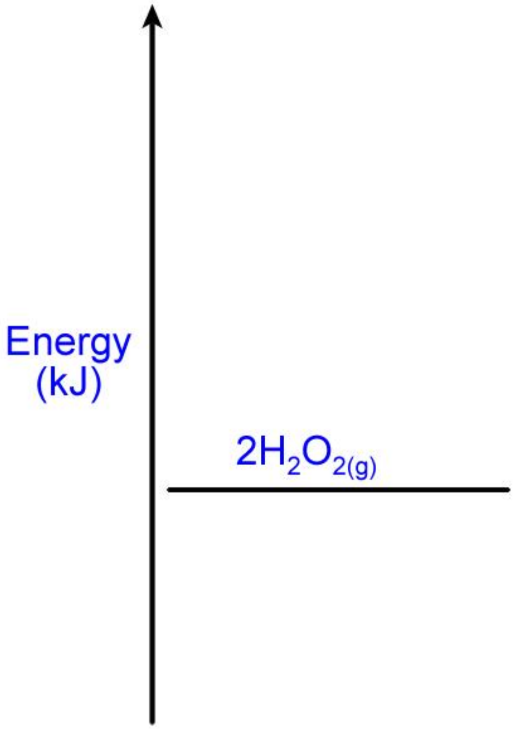
Energy Change Calculations – Example #2:
The Decomposition of Hydrogen Peroxide to form Oxygen and Water

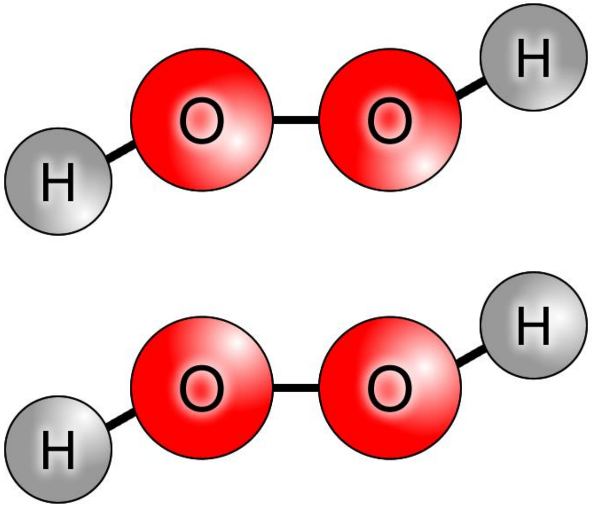
$2H_2O_{2(g)} \rightarrow O_{2(g)} + 2H_2O_{(g)}$

O–O = 146 kJ/mol

O–H = 463 kJ/mol

O=O = 496 kJ/mol





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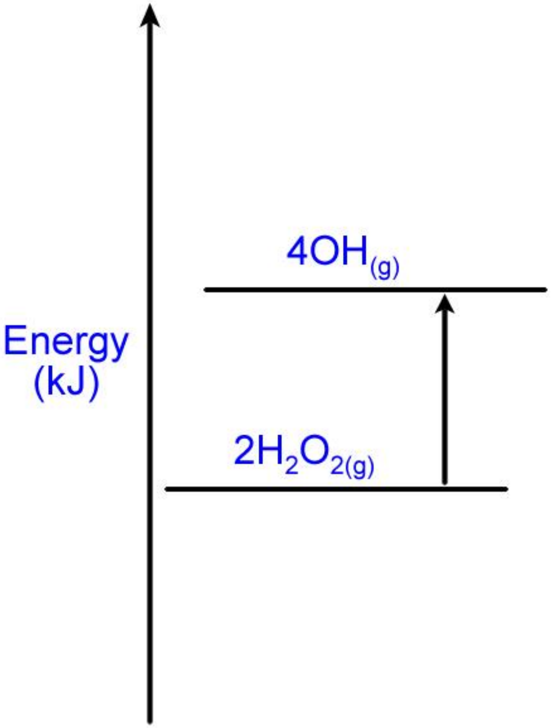

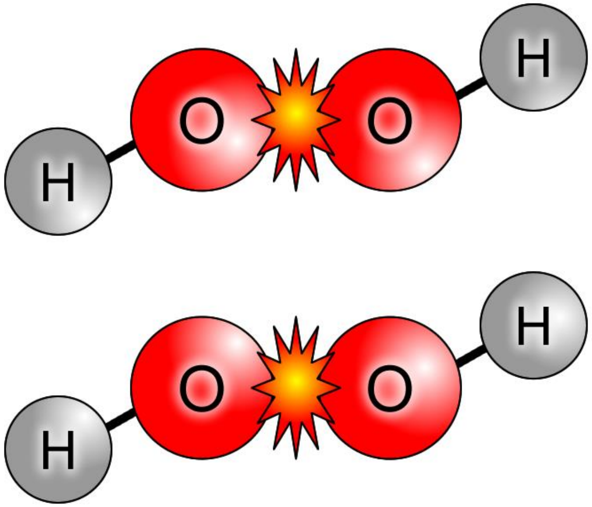

2 × O–O bond are broken forming 4 × OH molecules.

Energy Change Calculations – Example #2:

The Decomposition of Hydrogen Peroxide to form Oxygen and Water

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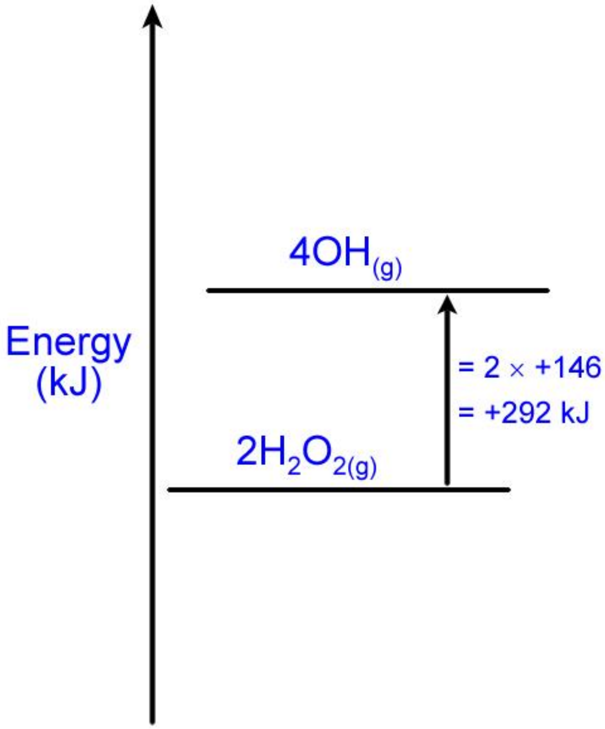
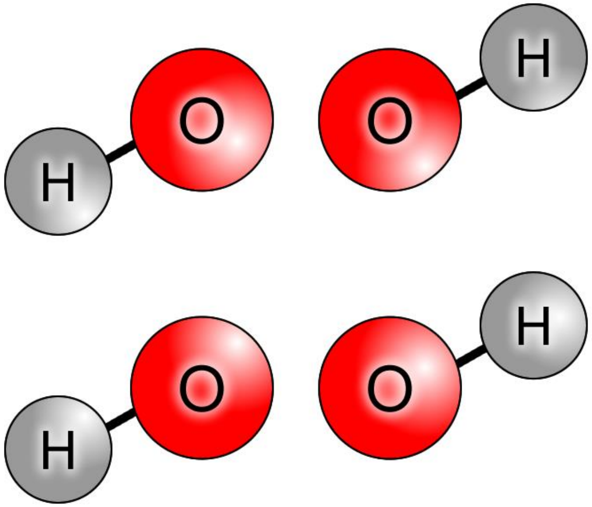
O–O = 146 kJ/mol O–H = 463 kJ/mol O=O = 496 kJ/mol





Bond breaking is *endothermic*: ΔH for this change is *positive*.

Energy Change Calculations – Example #2:

The Decomposition of Hydrogen Peroxide to form Oxygen and Water

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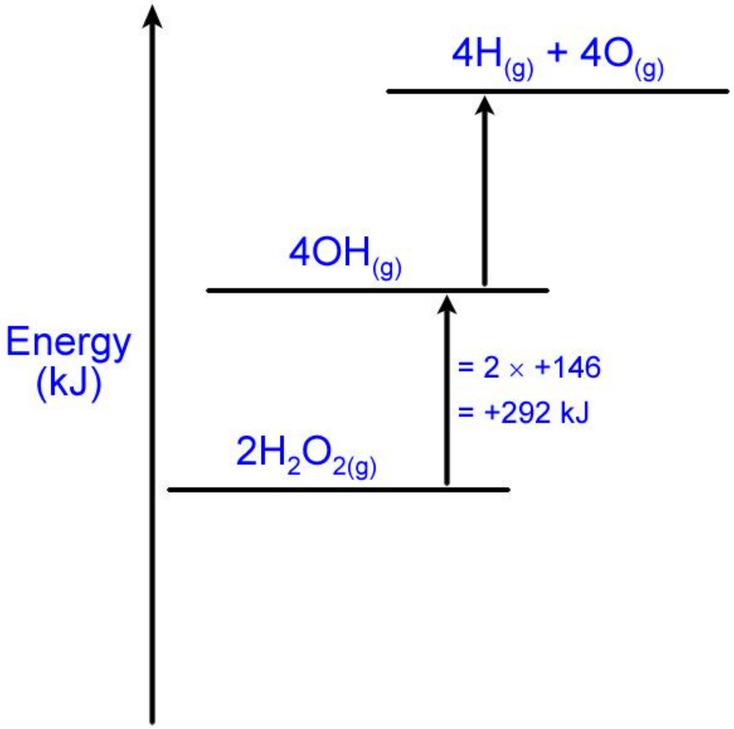

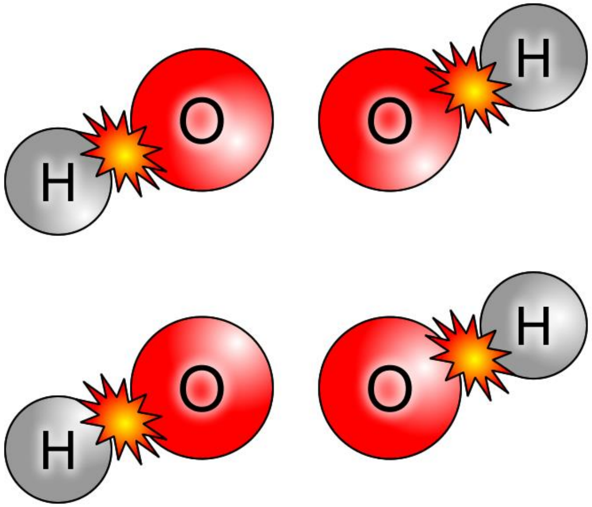

4 × O–H bonds are broken forming 4 × O and 4 × H atoms.

Energy Change Calculations – Example #2:

The Decomposition of Hydrogen Peroxide to form Oxygen and Water

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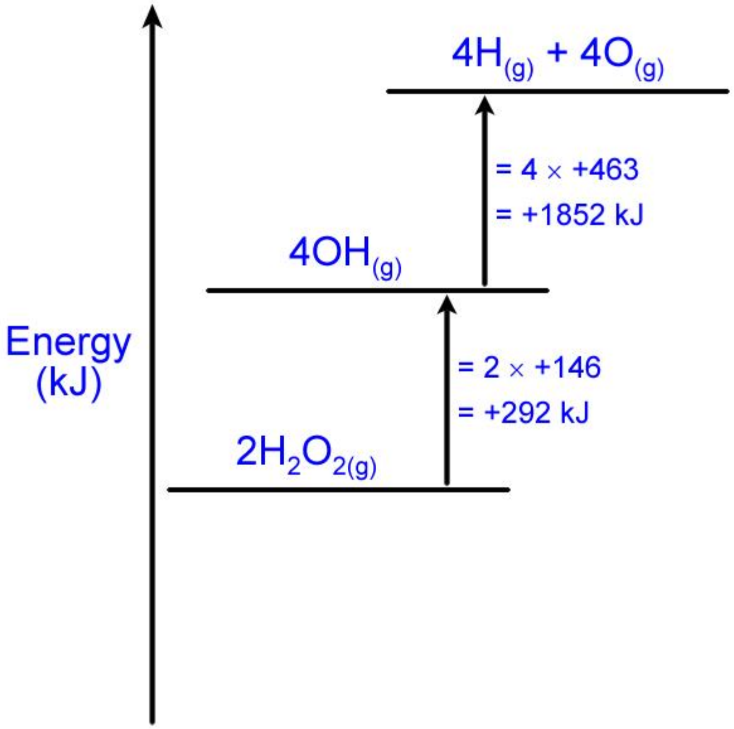

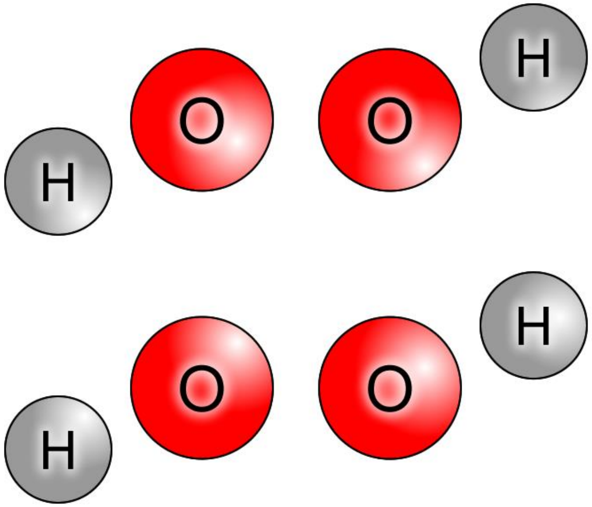

Bond breaking is *endothermic*: ΔH for this change is *positive*.

Energy Change Calculations – Example #2:

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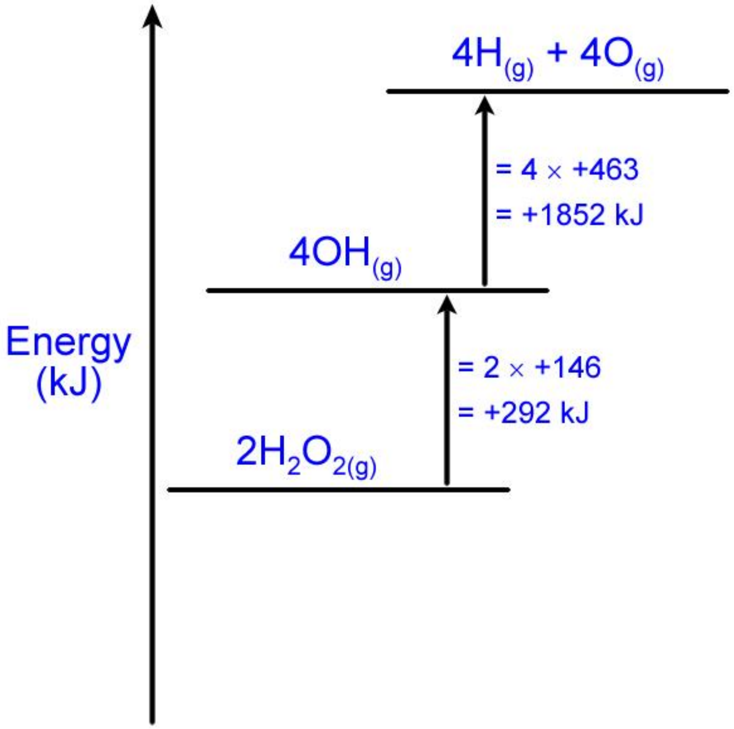

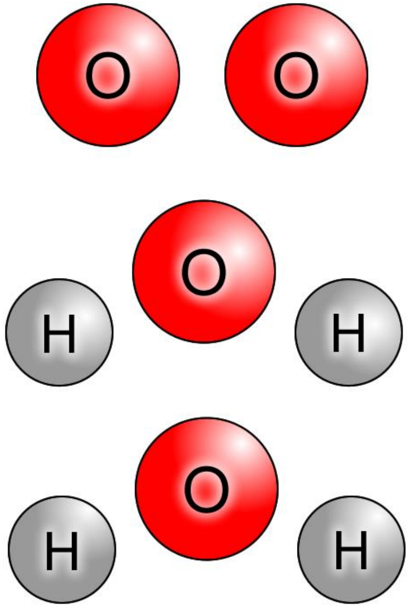

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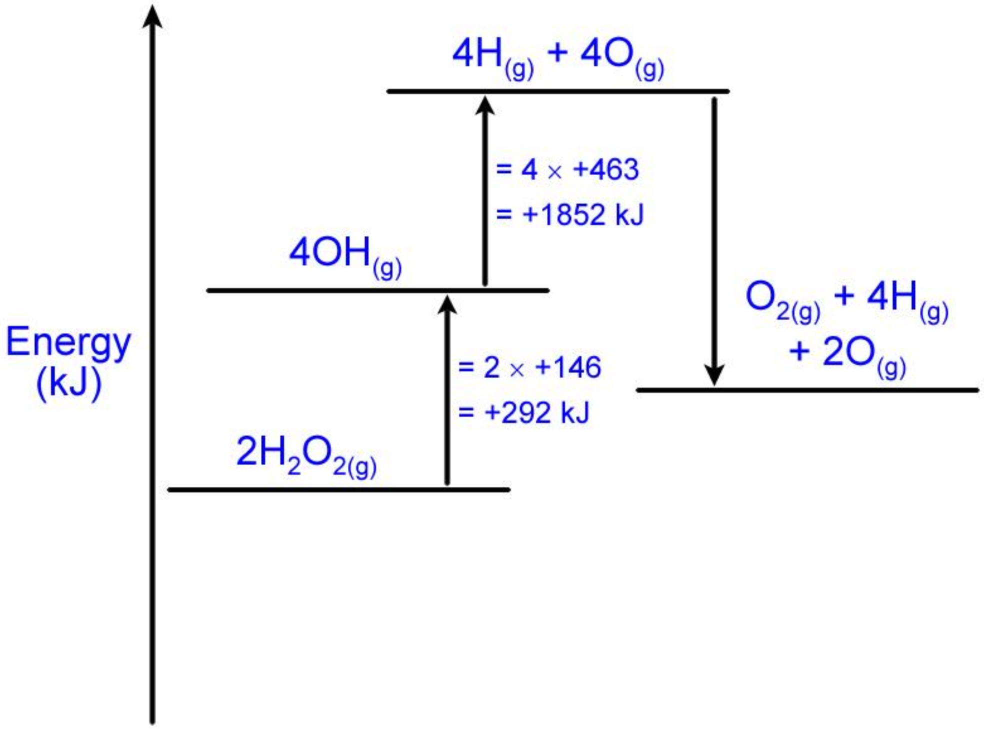

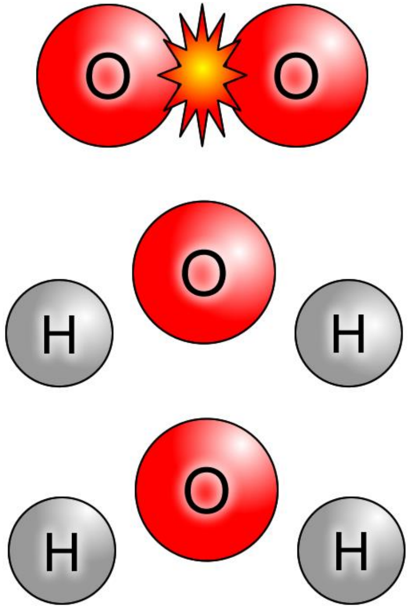

1 × O=O bond is formed creating 1 × O_2 molecule.

Energy Change Calculations – Example #2:

The Decomposition of Hydrogen Peroxide to form Oxygen and Water

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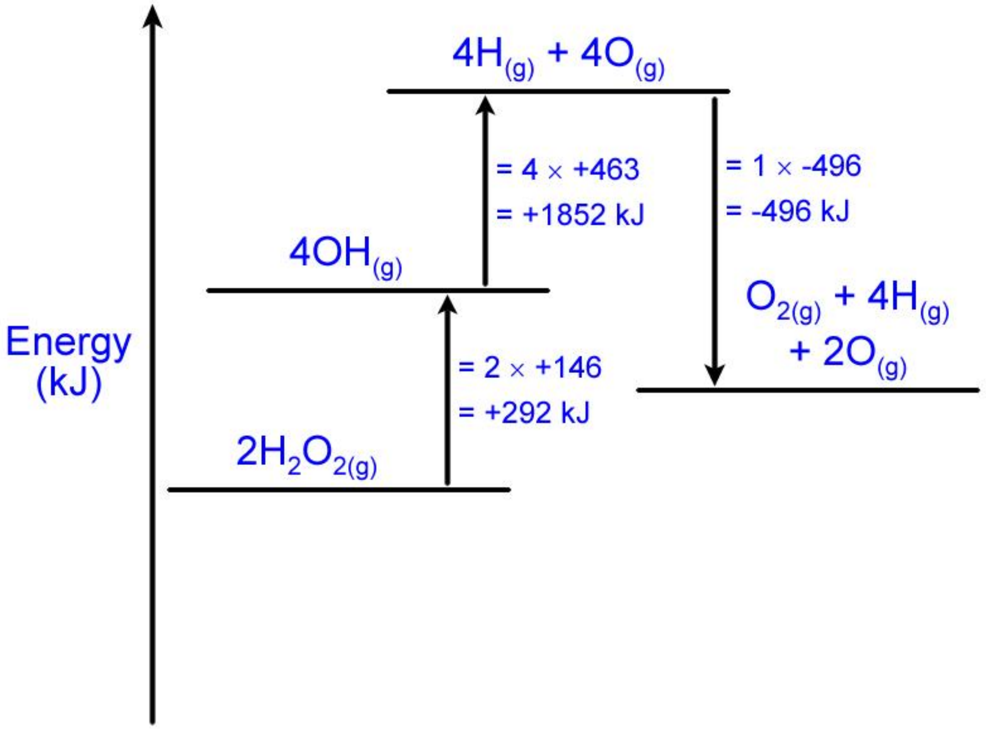

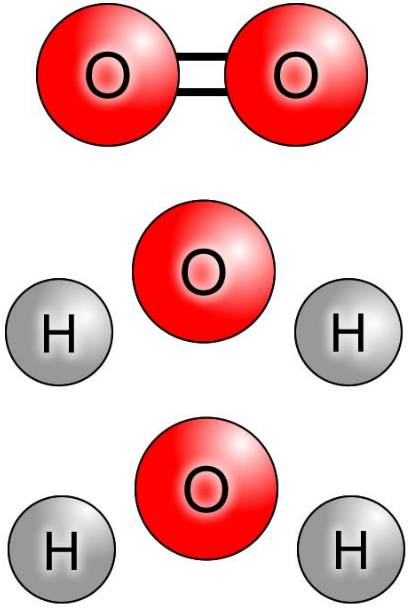

Bond formation is *exothermic*: ΔH for this change is *negative*.

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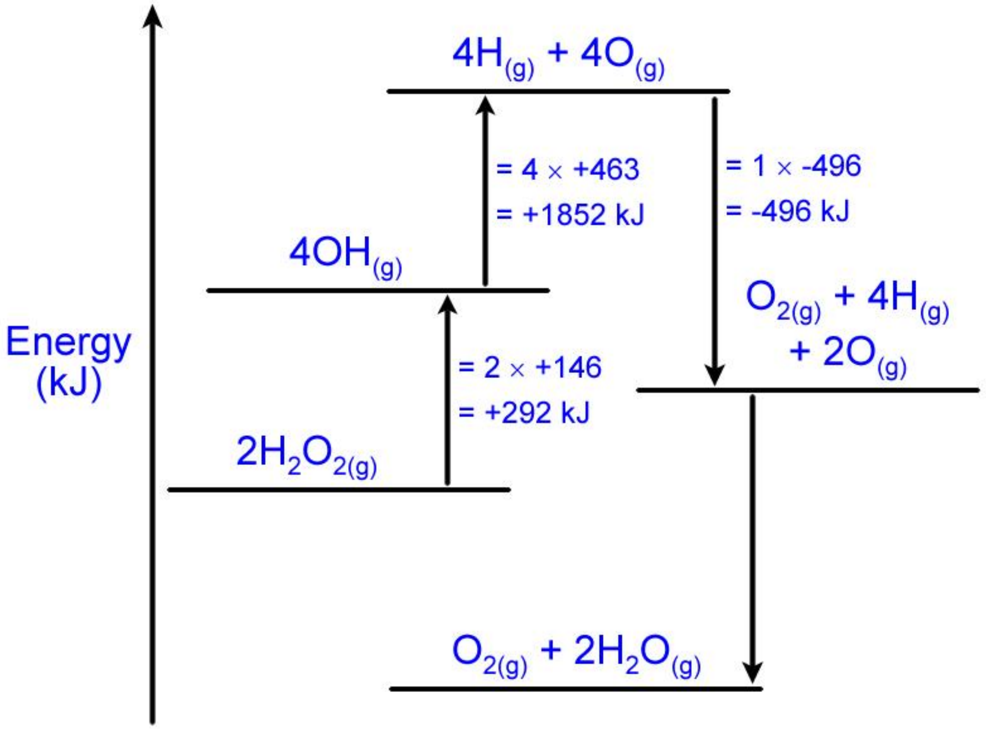
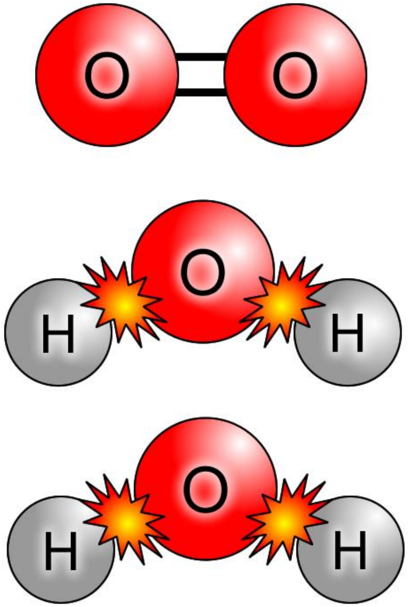




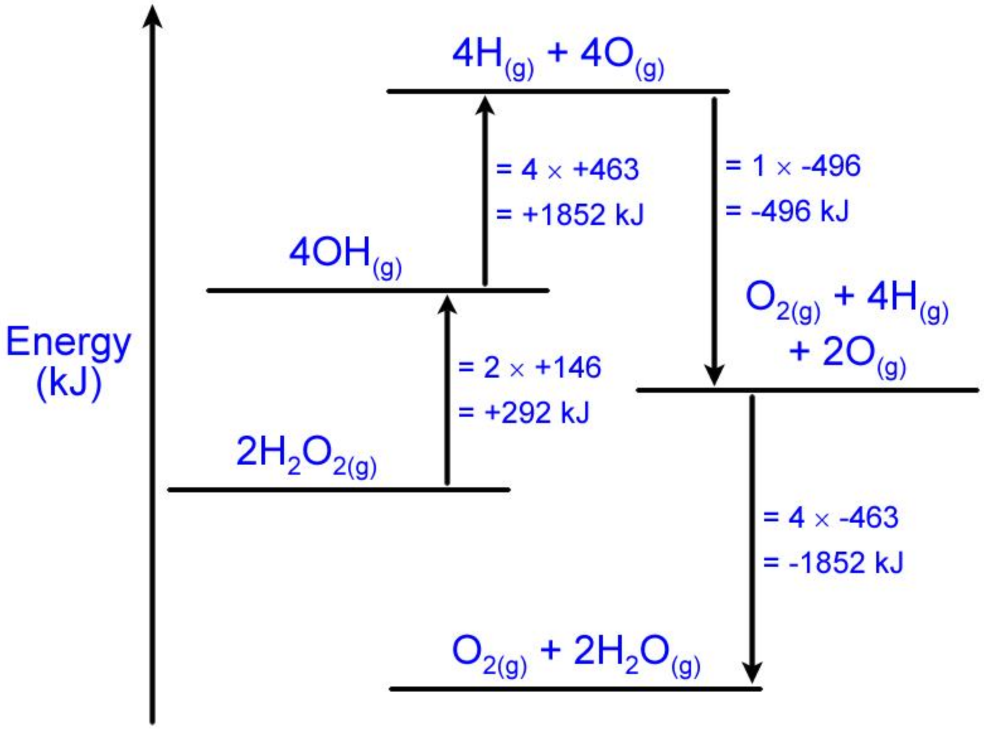
4 × O–H bonds are formed creating 2 × H_2O molecules.

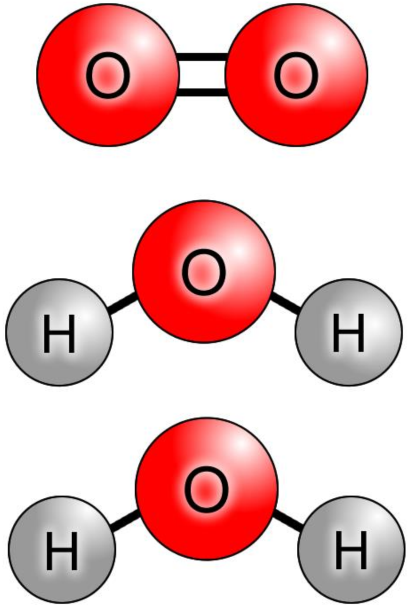
Energy Change Calculations – Example #2:

The Decomposition of Hydrogen Peroxide to form Oxygen and Water

$$2H_2O_{2(g)} \rightarrow O_{2(g)} + 2H_2O_{(g)}$$

O–O = 146 kJ/mol   O–H = 463 kJ/mol   O=O = 496 kJ/mol





Bond formation is *exothermic*: ΔH for this change is *negative*.

Energy Change Calculations – Example #2:

The Decomposition of Hydrogen Peroxide to form Oxygen and Water

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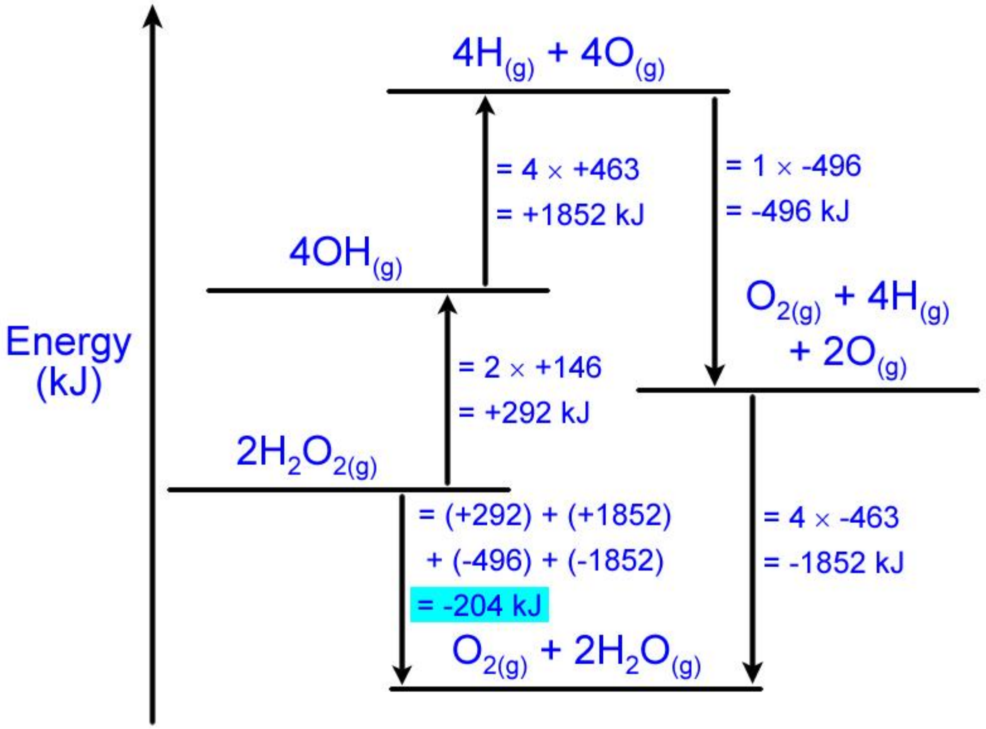


Energy (kJ)

$4H_{(g)} + 4O_{(g)}$

$= 4 \times +463$
$= +1852$ kJ

$= 1 \times -496$
$= -496$ kJ

$4OH_{(g)}$

$O_{2(g)} + 4H_{(g)} + 2O_{(g)}$

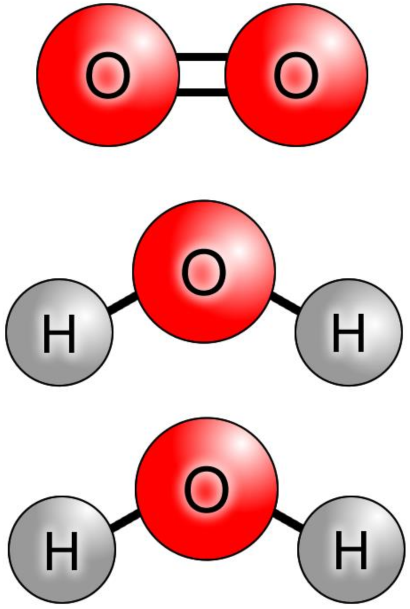
$= 2 \times +146$
$= +292$ kJ

$2H_2O_{2(g)}$

= (+292) + (+1852) + (-496) + (-1852)
= -204 kJ

$= 4 \times -463$
$= -1852$ kJ

$O_{2(g)} + 2H_2O_{(g)}$



$$(+292) + (+1852) + (-496) + (-1852) = -204 \text{ kJ}$$

Energy Change Calculations – Example #2:

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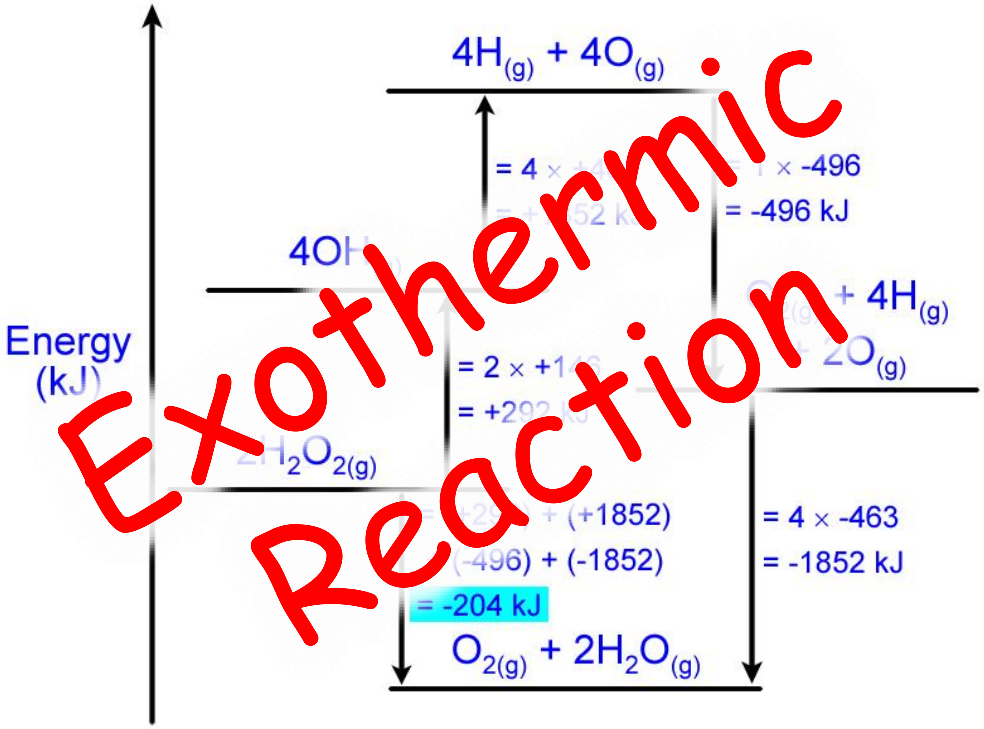
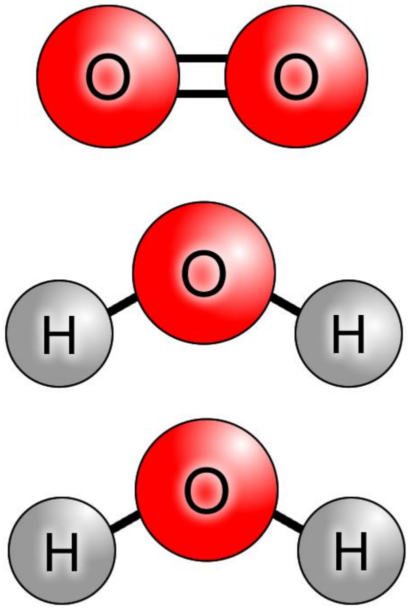


Exothermic Reaction

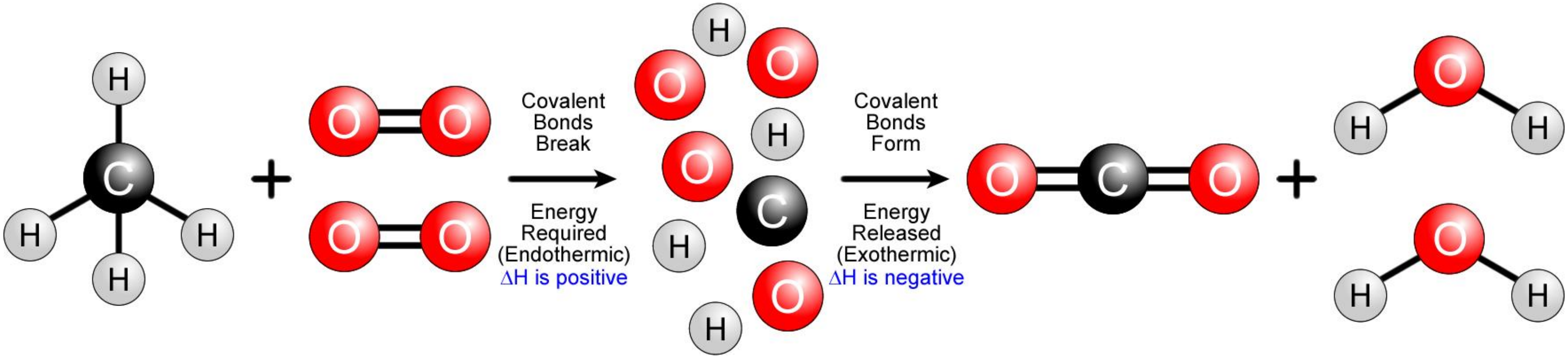


(+292) + (+1852) + (−496) + (−1852) = *−204 kJ*

Energy Change Calculations – Example #3:
The Combustion of Methane forming Carbon Dioxide and Water

$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)}$

Overview of the Reaction:



Covalent Bonds Break

Energy Required (Endothermic)

ΔH is positive

Covalent Bonds Form

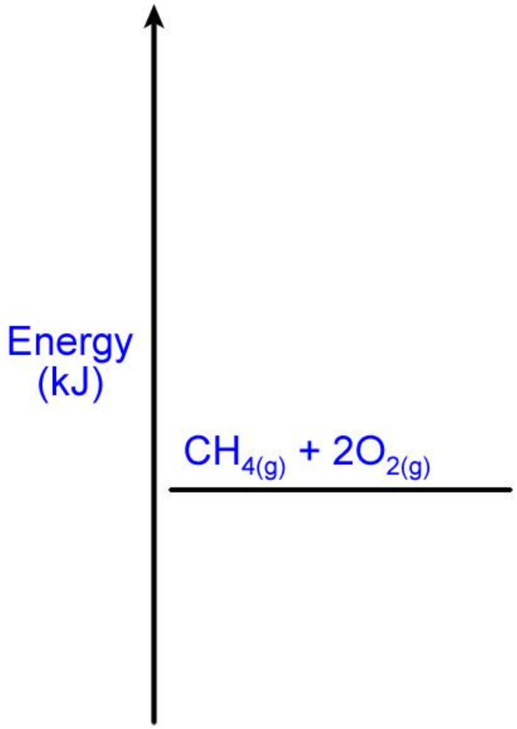
Energy Released (Exothermic)

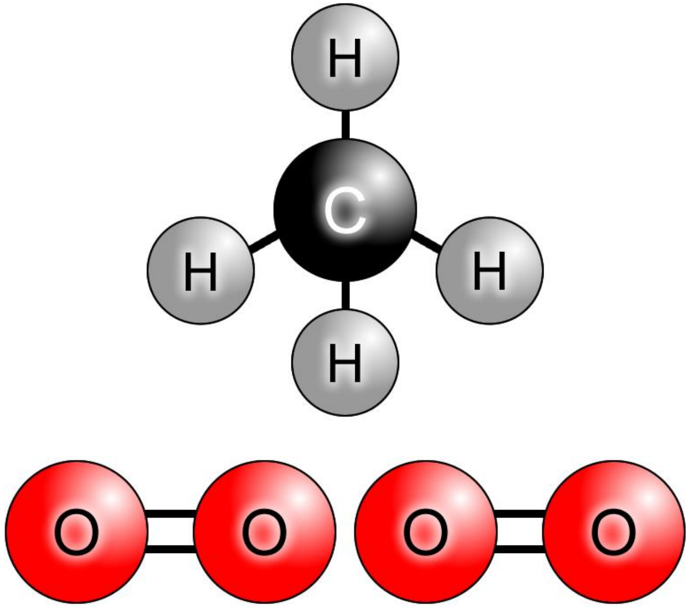
ΔH is negative

Energy Change Calculations – Example #3:
The Combustion of Methane forming Carbon Dioxide and Water

$$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)}$$

C–H = 412 kJ/mol O=O = 496 kJ/mol C=O = 743 kJ/mol O–H = 463 kJ/mol





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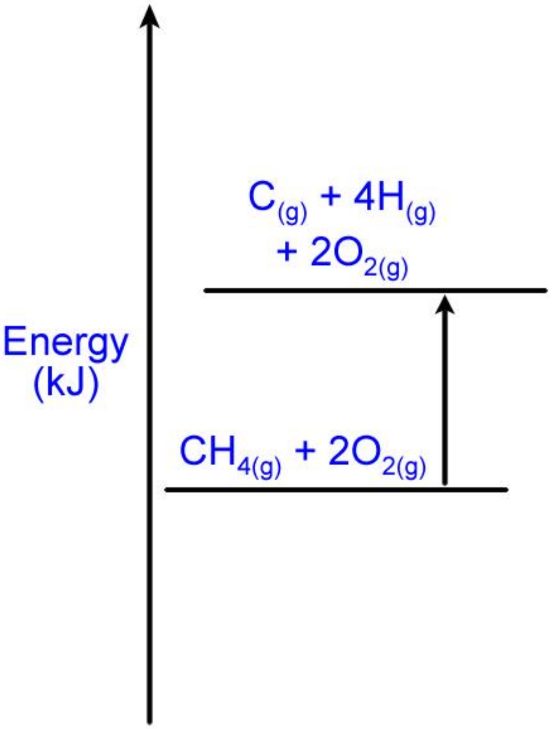

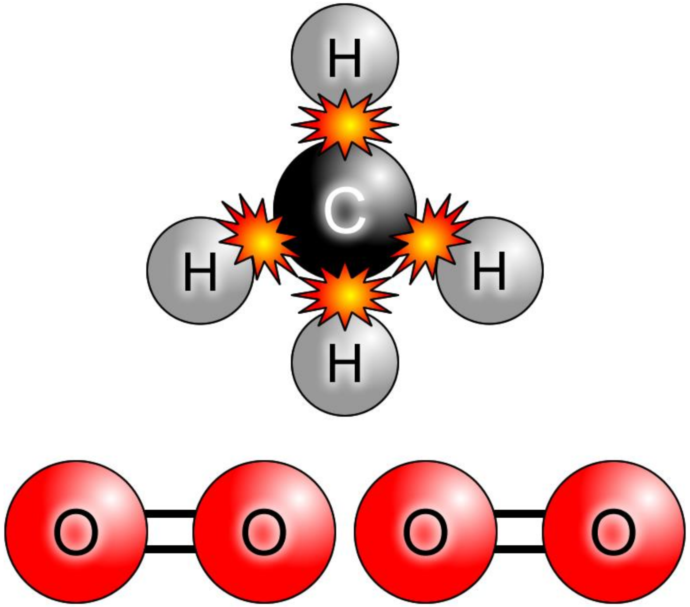

4 × C–H bonds are broken forming 1 × C atom and 4 × H atoms.

Energy Change Calculations – Example #3:

The Combustion of Methane forming Carbon Dioxide and Water

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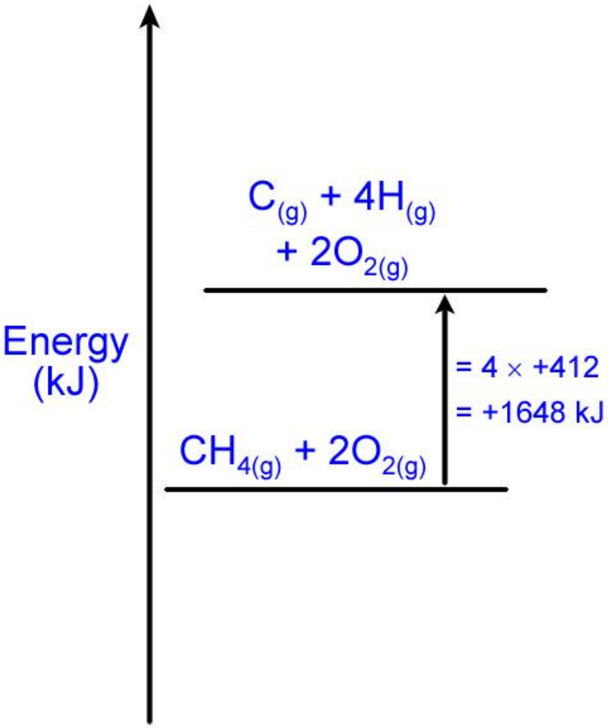

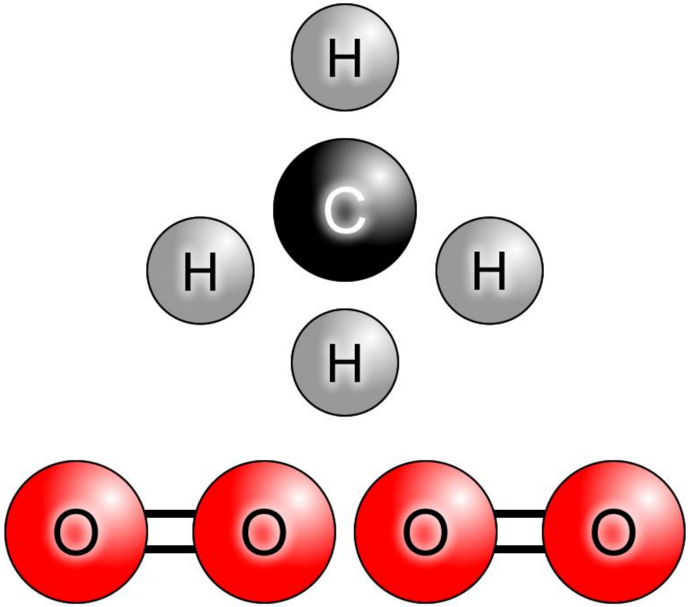

Bond breaking is *endothermic*: ΔH for this change is *positive*.

Energy Change Calculations – Example #3:

The Combustion of Methane forming Carbon Dioxide and Water

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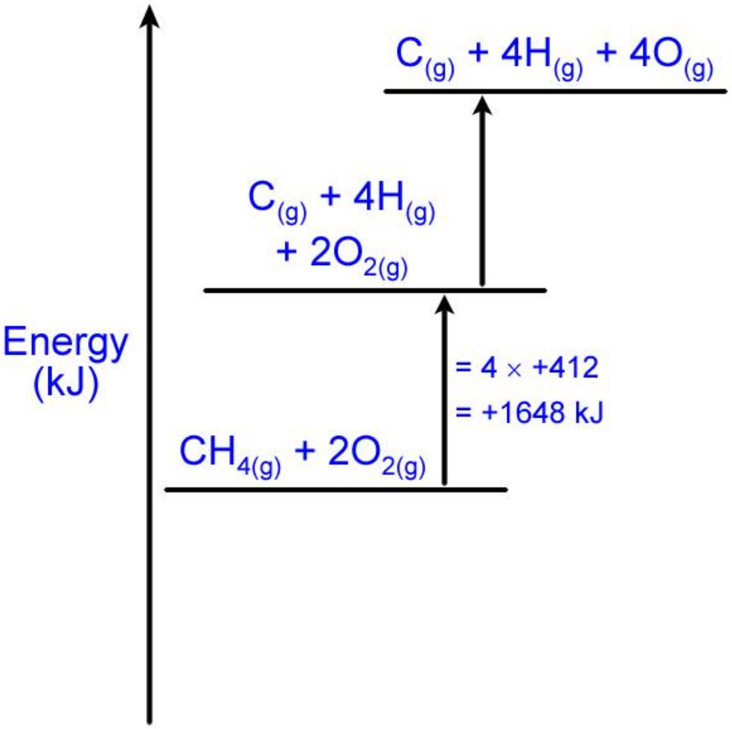


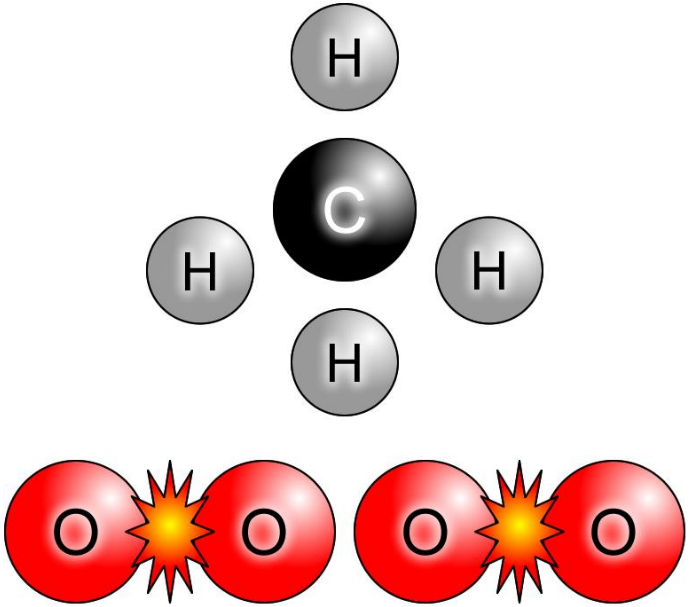
$C_{(g)} + 4H_{(g)} + 4O_{(g)}$

$C_{(g)} + 4H_{(g)}$
$+ 2O_{2(g)}$

Energy
(kJ)

= 4 × +412
= +1648 kJ

$CH_{4(g)} + 2O_{2(g)}$



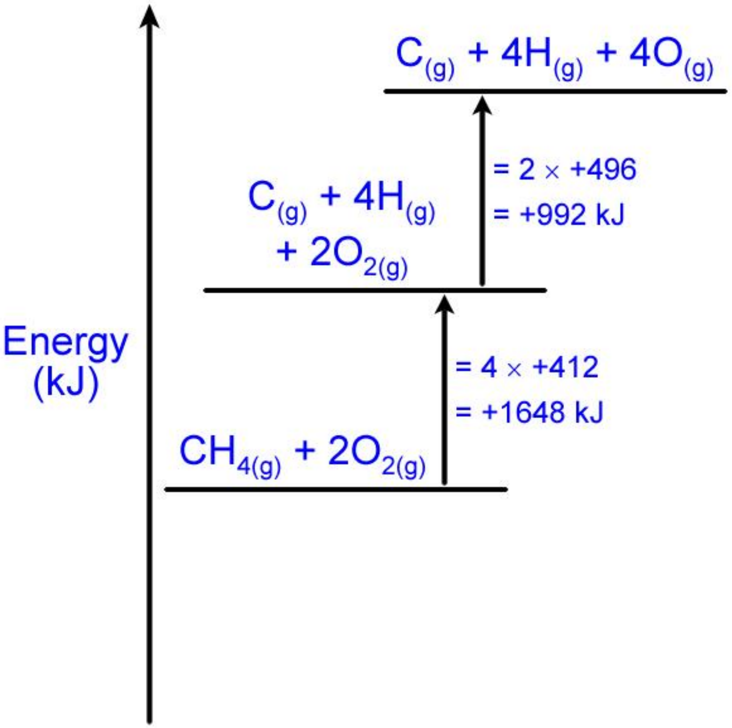
2 × O=O bonds are broken forming 4 × O atoms

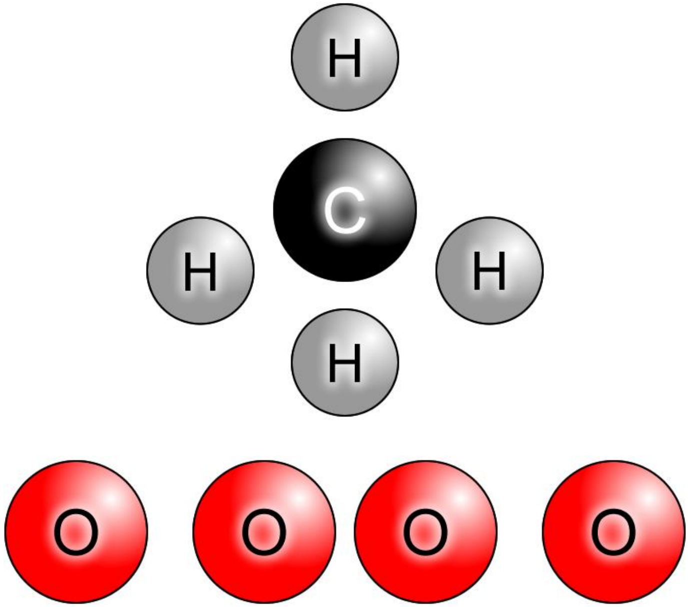
Energy Change Calculations – Example #3:

The Combustion of Methane forming Carbon Dioxide and Water

$$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)}$$

C–H = 412 kJ/mol   O=O = 496 kJ/mol   C=O = 743 kJ/mol   O–H = 463 kJ/mol





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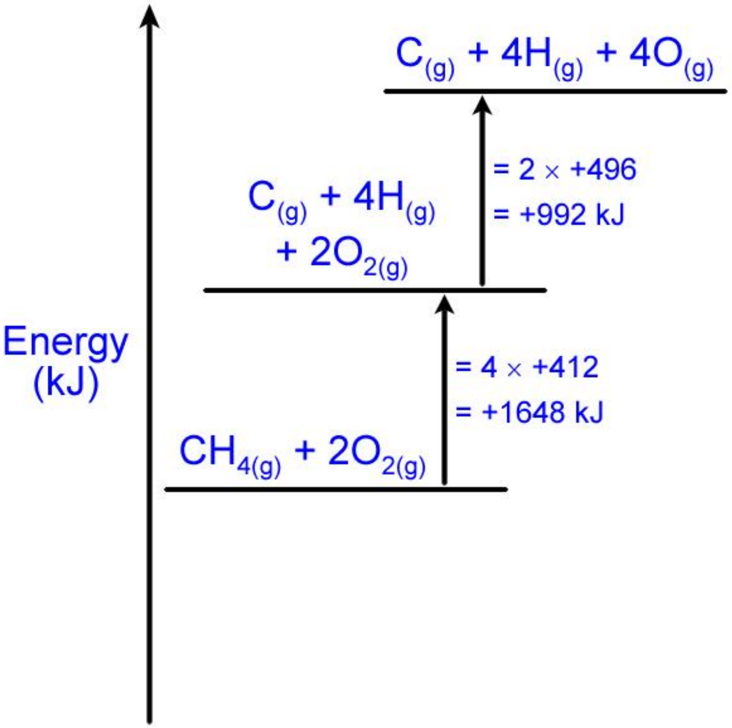

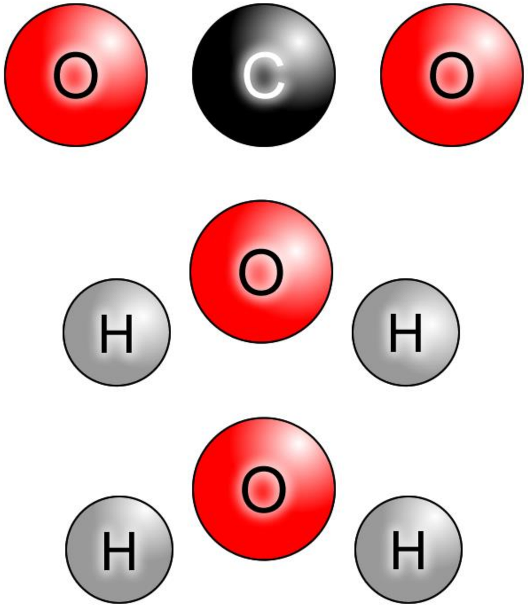

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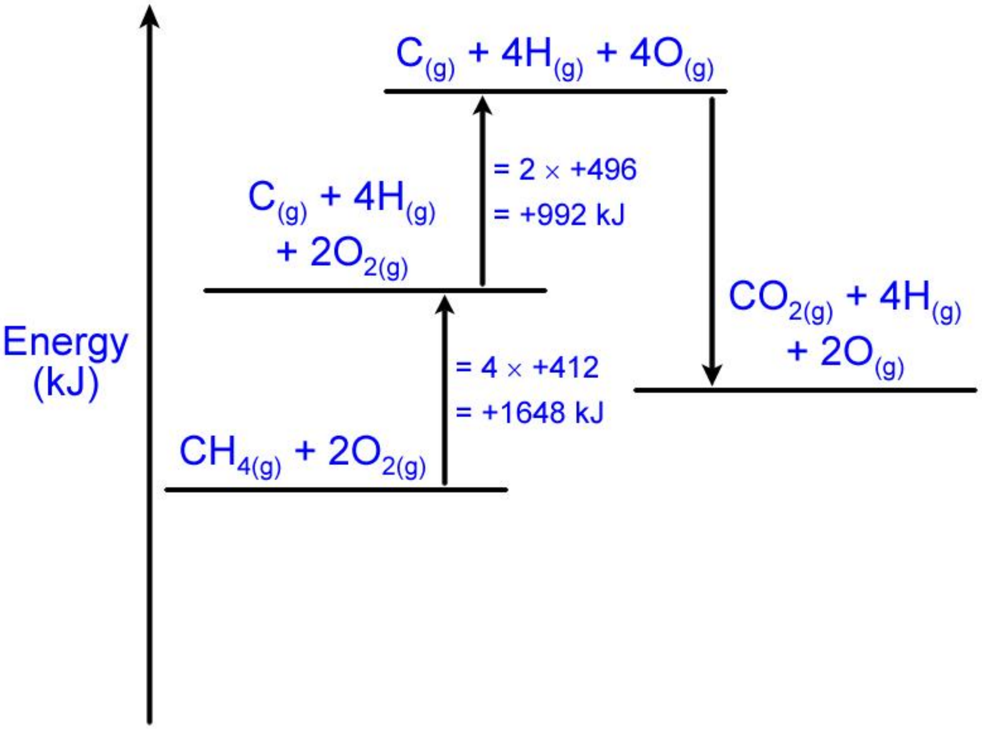

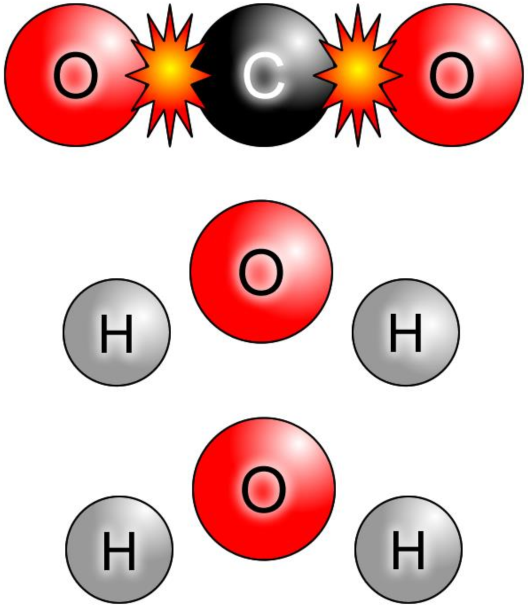

2 × C=O bonds are formed creating 1 × CO_2 molecule.

Energy Change Calculations – Example #3:

The Combustion of Methane forming Carbon Dioxide and Water

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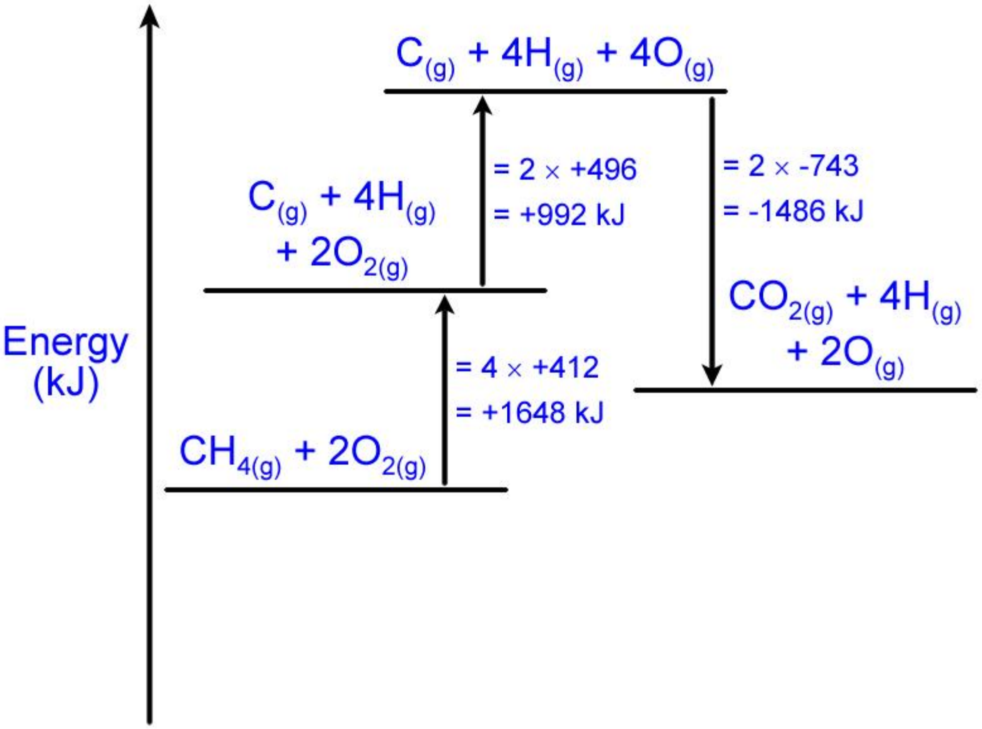


Energy (kJ)

$C_{(g)} + 4H_{(g)} + 4O_{(g)}$

$C_{(g)} + 4H_{(g)} + 2O_{2(g)}$

= 2 × +496

= +992 kJ

= 2 × -743

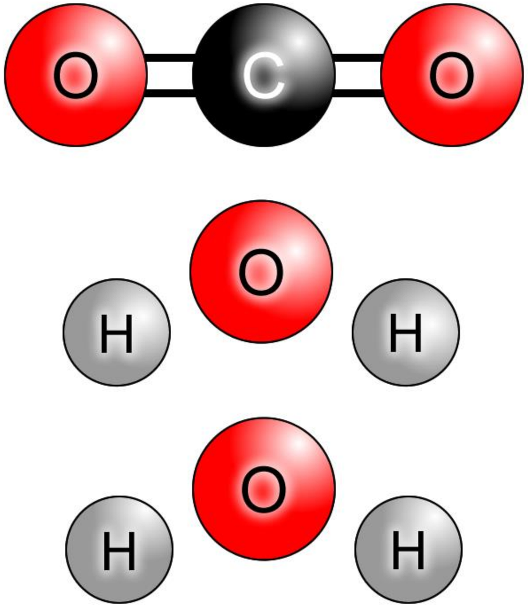
= -1486 kJ

$CO_{2(g)} + 4H_{(g)} + 2O_{(g)}$

= 4 × +412

= +1648 kJ

$CH_{4(g)} + 2O_{2(g)}$



Bond formation is *exothermic*: ΔH for this change is *negative*.

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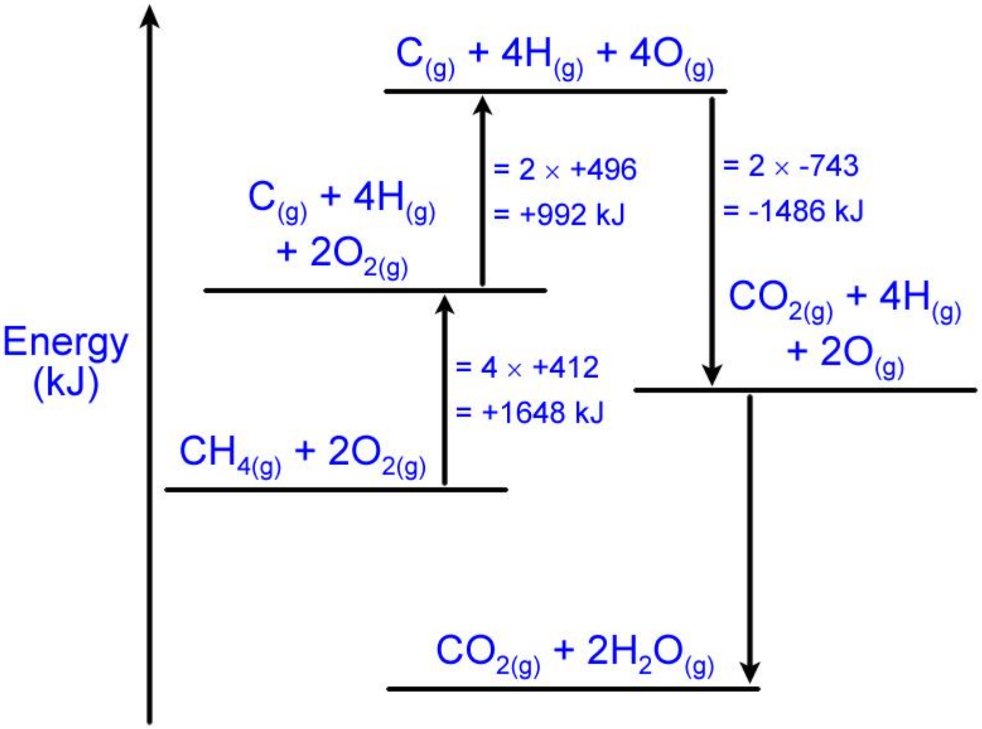

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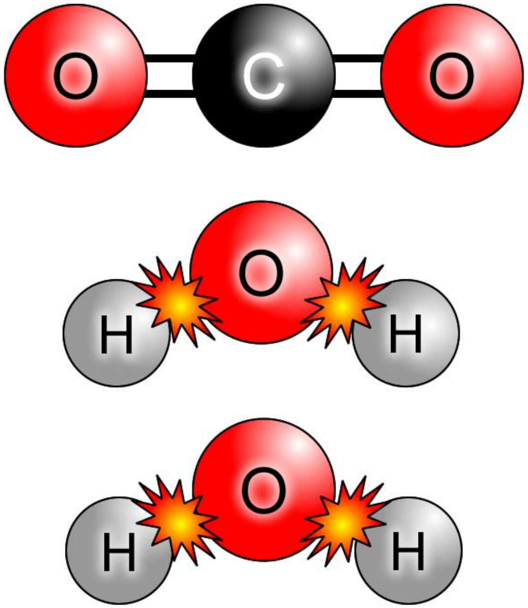

4 × O–H bonds are formed creating 2 × H_2O molecules.

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The Combustion of Methane forming Carbon Dioxide and Water

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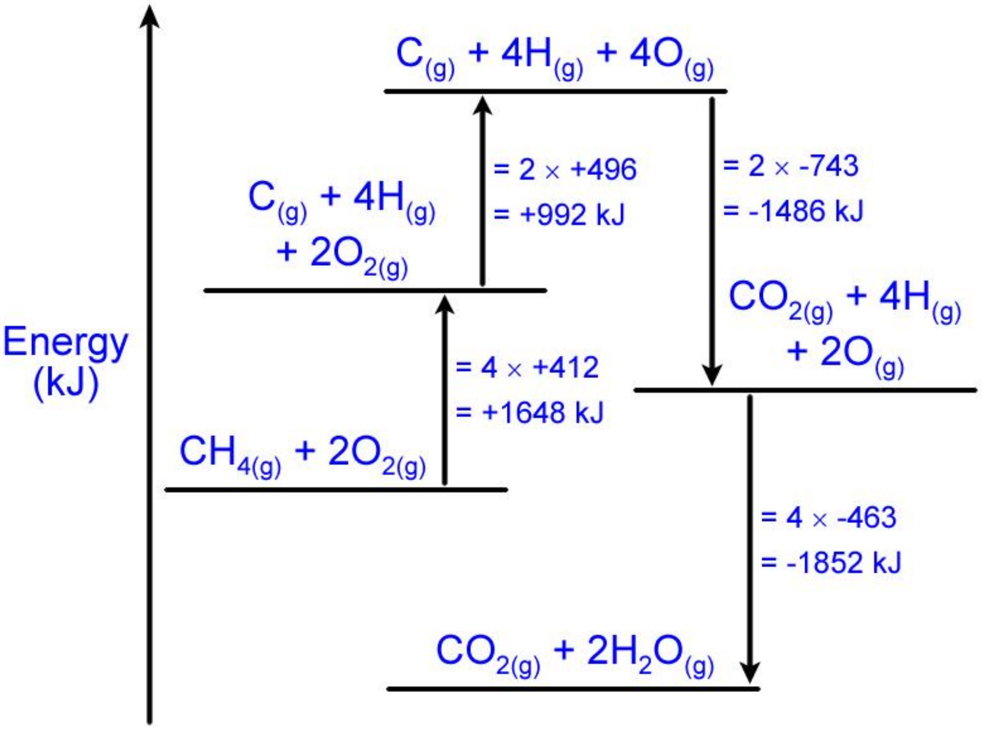

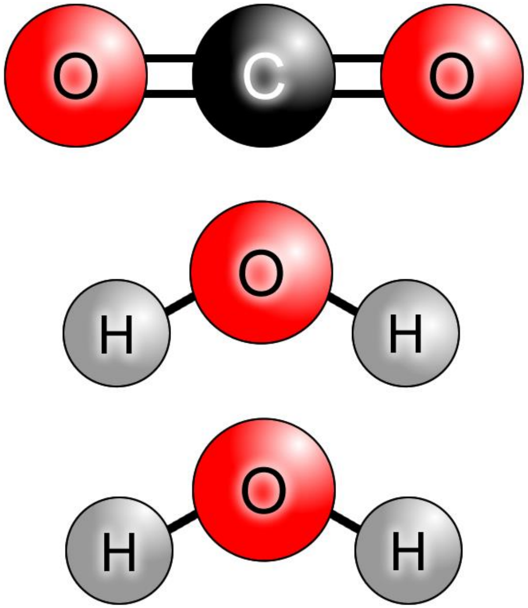

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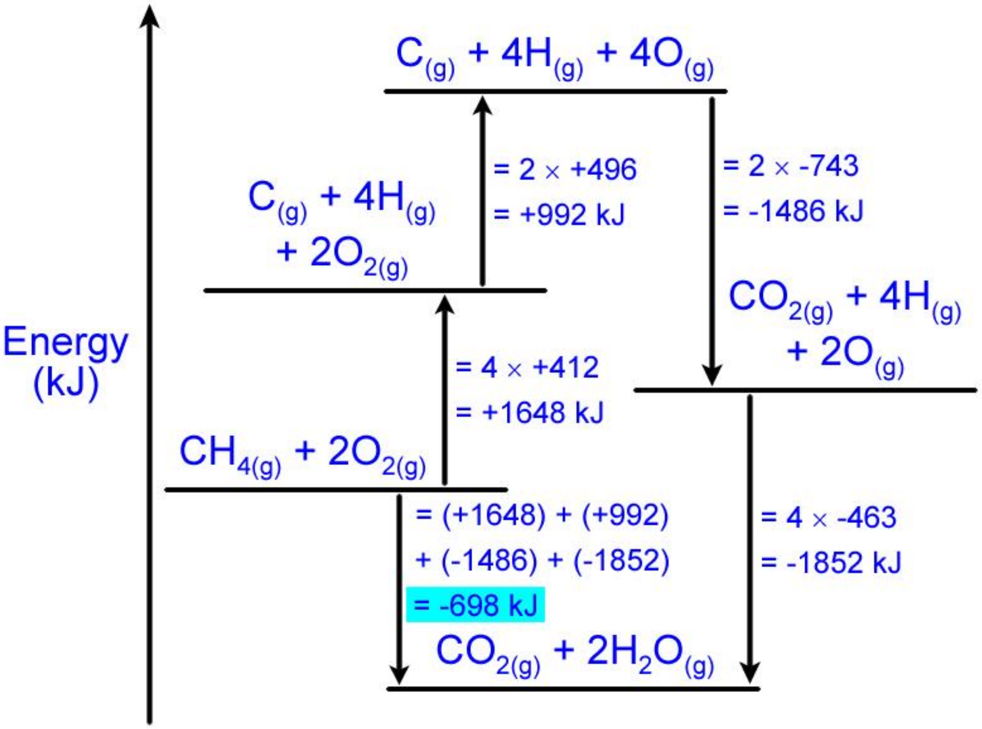

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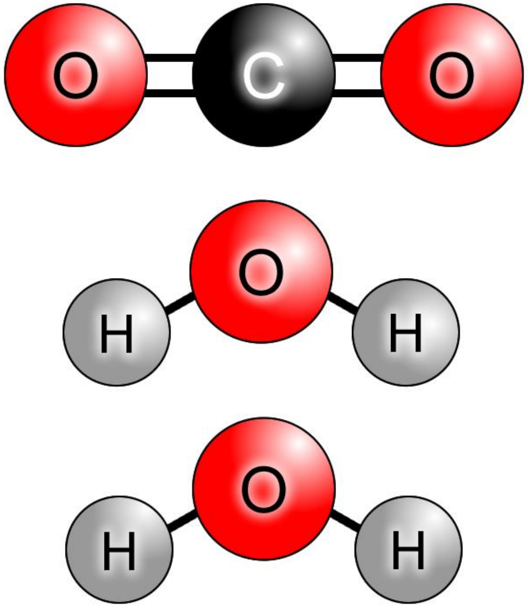
= 4 × +412
= +1648 kJ

$CH_{4(g)} + 2O_{2(g)}$

= (+1648) + (+992)
+ (-1486) + (-1852)
= -698 kJ

= 4 × -463
= -1852 kJ

$CO_{2(g)} + 2H_2O_{(g)}$



(+1648) + (+992) + (−1486) + (−1852) = *−698 kJ*

Energy Change Calculations – Example #3:

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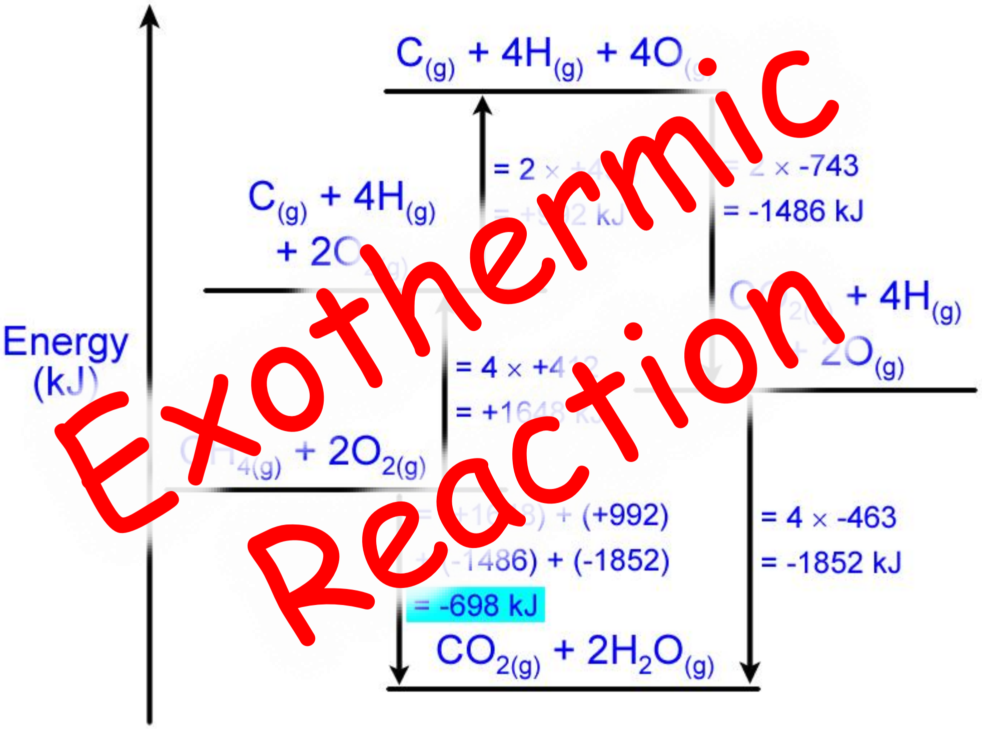
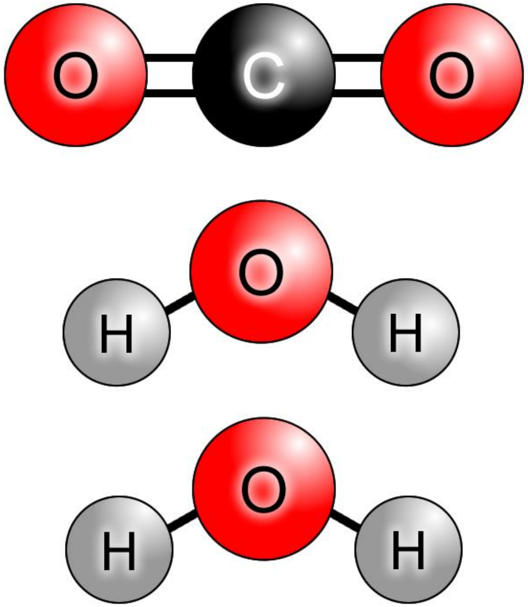


Exothermic Reaction



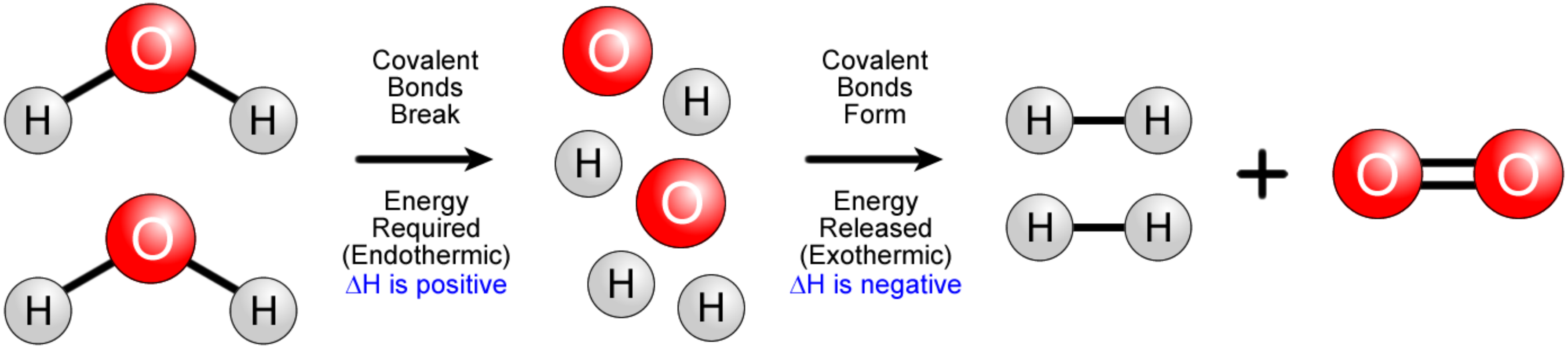
$$(+1648) + (+992) + (-1486) + (-1852) = \boxed{-698\ kJ}$$

Energy Change Calculations – Example #4:

The Electrolysis of Water to form Hydrogen and Oxygen

$$2H_2O_{(g)} \rightarrow 2H_{2(g)} + O_{2(g)}$$

Overview of the Reaction:



Covalent Bonds Break

Energy Required (Endothermic)

ΔH is positive

Covalent Bonds Form

Energy Released (Exothermic)

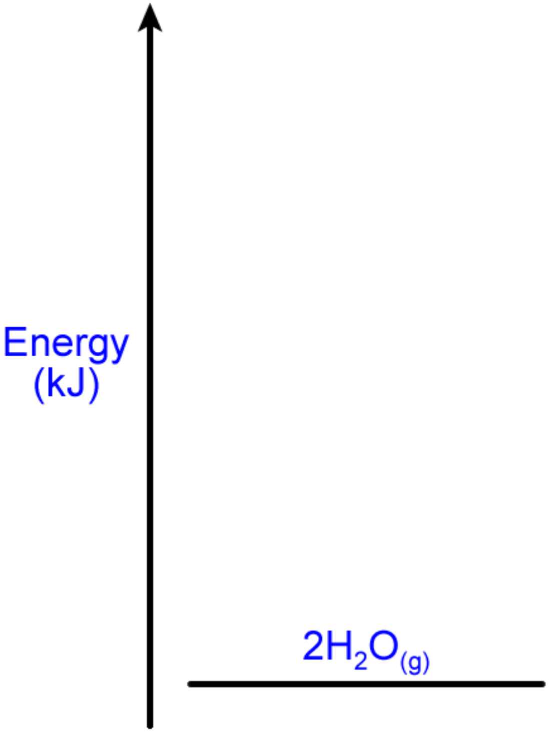
ΔH is negative

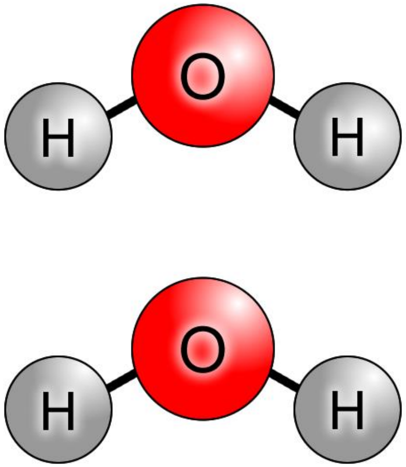
Energy Change Calculations – Example #4:

The Electrolysis of Water to form Hydrogen and Oxygen

$$2H_2O_{(g)} \rightarrow 2H_{2(g)} + O_{2(g)}$$

O–H = 463 kJ/mol   O=O = 496 kJ/mol   H–H = 436 kJ/mol





Energy Change Calculations – Example #4:

The Electrolysis of Water to form Hydrogen and Oxygen

$$2H_2O_{(g)} \rightarrow 2H_{2(g)} + O_{2(g)}$$

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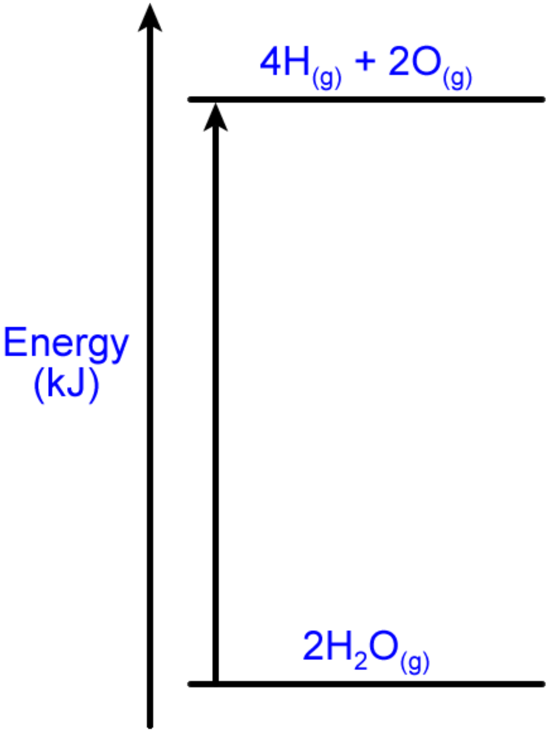
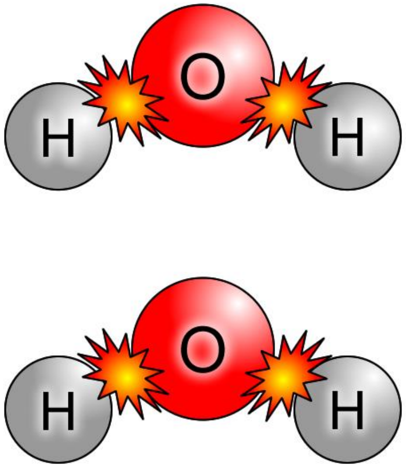




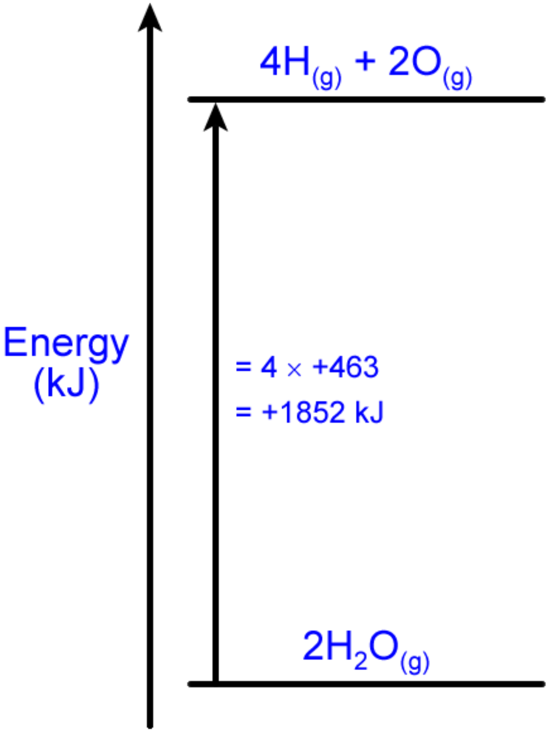
4 × O–H bonds are broken forming 2 × O and 4 × H atoms.

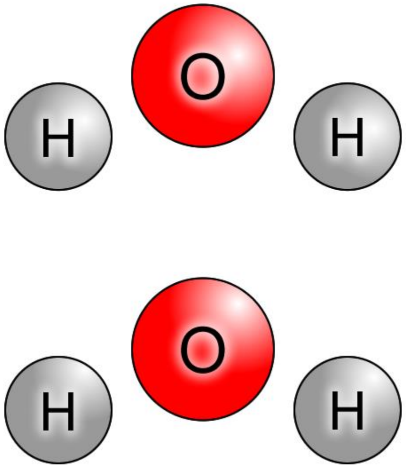
Energy Change Calculations – Example #4:

The Electrolysis of Water to form Hydrogen and Oxygen

$$2H_2O_{(g)} \rightarrow 2H_{2(g)} + O_{2(g)}$$

O–H = 463 kJ/mol   O=O = 496 kJ/mol   H–H = 436 kJ/mol





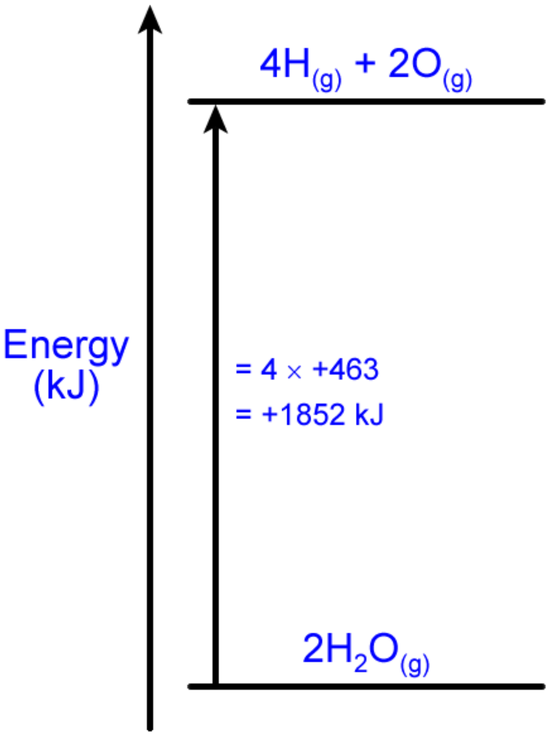
Bond breaking is *endothermic*: ΔH for this change is *positive*.

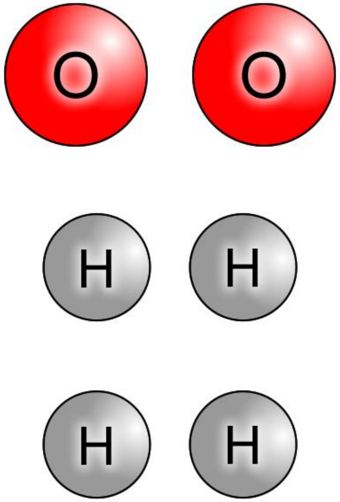
Energy Change Calculations – Example #4:

The Electrolysis of Water to form Hydrogen and Oxygen

$$2H_2O_{(g)} \rightarrow 2H_{2(g)} + O_{2(g)}$$

O–H = 463 kJ/mol   O=O = 496 kJ/mol   H–H = 436 kJ/mol





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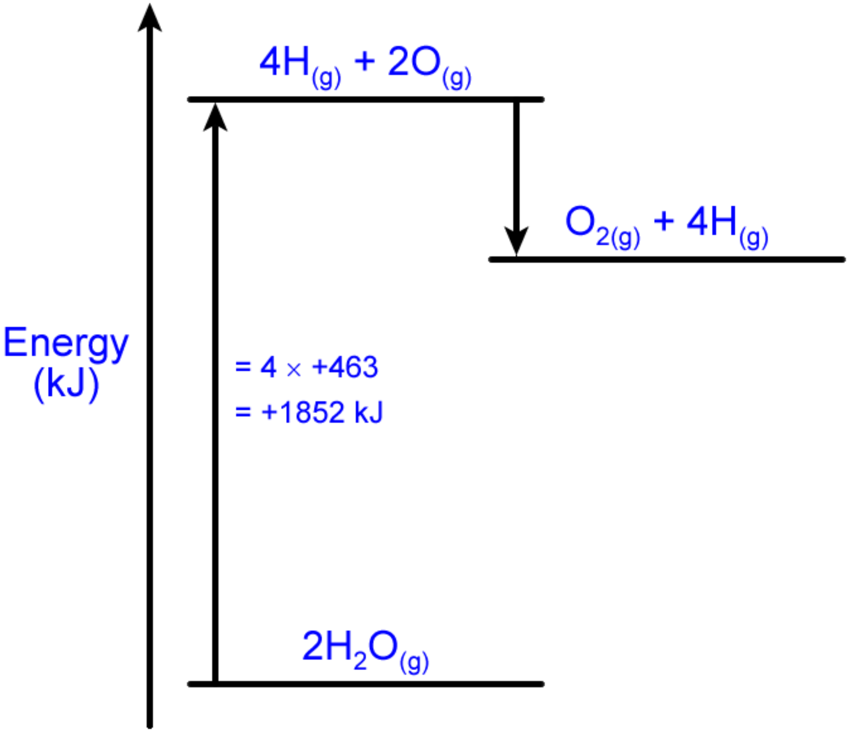


$4H_{(g)} + 2O_{(g)}$

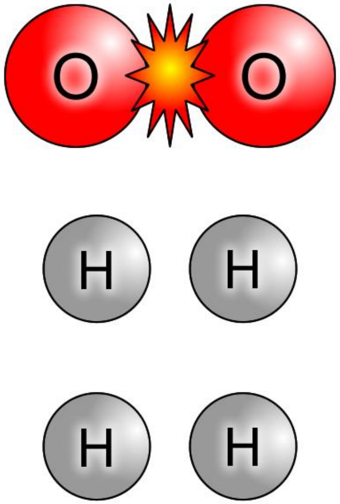
$O_{2(g)} + 4H_{(g)}$

Energy (kJ)

$= 4 \times +463$

$= +1852$ kJ

$2H_2O_{(g)}$



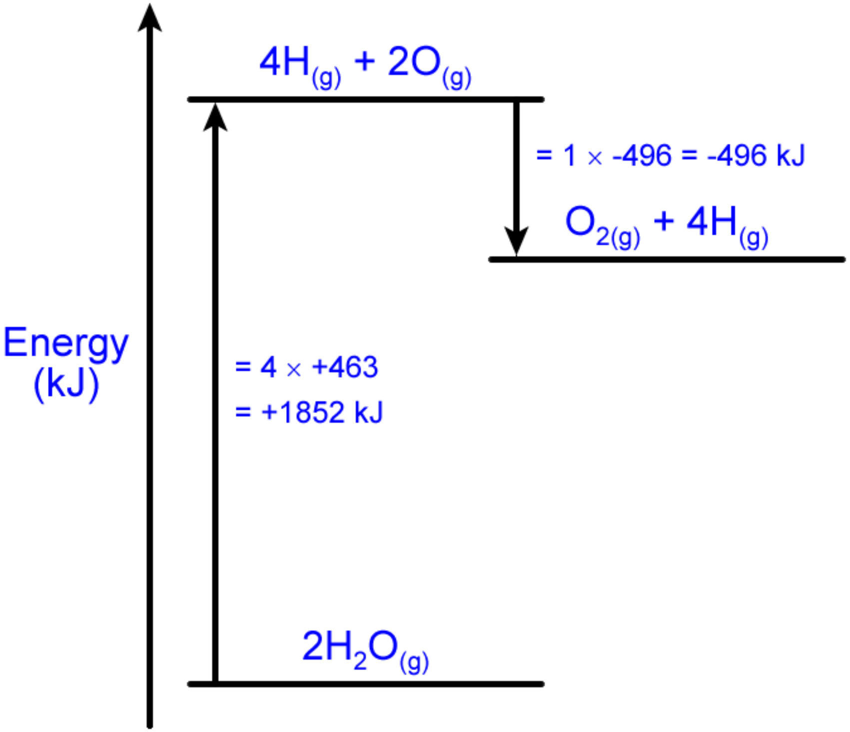
1 × O=O bond is formed creating 1 × O_2 molecule.

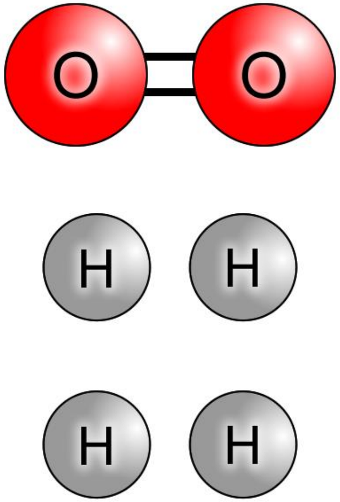
Energy Change Calculations – Example #4:

The Electrolysis of Water to form Hydrogen and Oxygen

$$2H_2O_{(g)} \rightarrow 2H_{2(g)} + O_{2(g)}$$

O–H = 463 kJ/mol   O=O = 496 kJ/mol   H–H = 436 kJ/mol





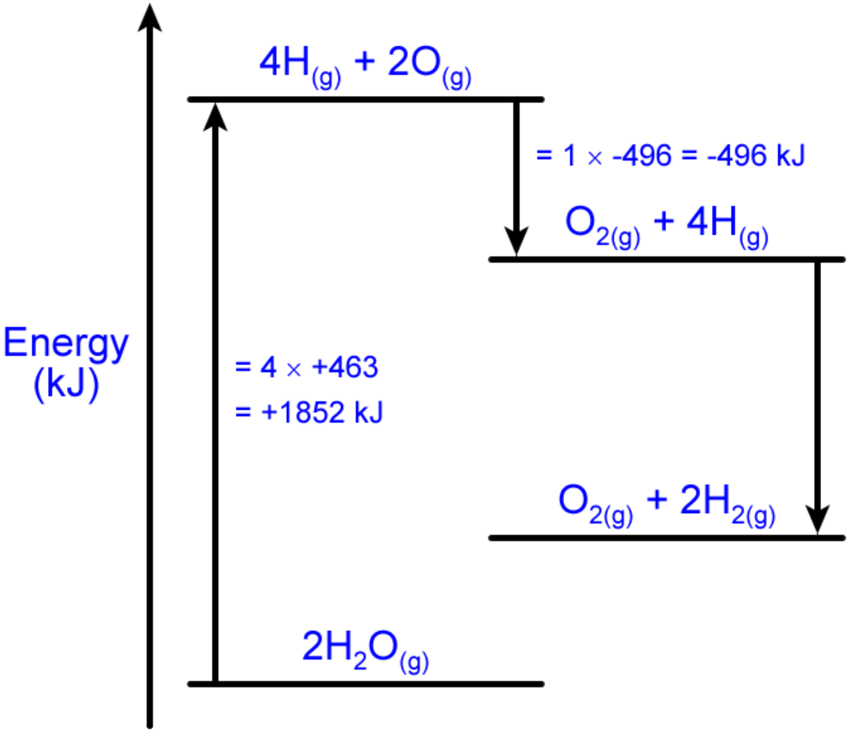
Bond formation is *exothermic*: ΔH for this change is *negative*.

Energy Change Calculations – Example #4:

The Electrolysis of Water to form Hydrogen and Oxygen

$$2H_2O_{(g)} \rightarrow 2H_{2(g)} + O_{2(g)}$$

O–H = 463 kJ/mol O=O = 496 kJ/mol H–H = 436 kJ/mol



Energy (kJ)

$4H_{(g)} + 2O_{(g)}$

= 1 × -496 = -496 kJ

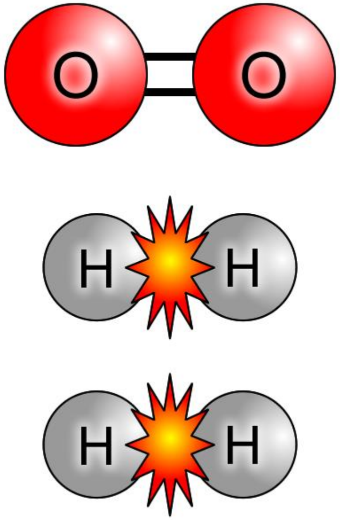
$O_{2(g)} + 4H_{(g)}$

= 4 × +463

= +1852 kJ

$O_{2(g)} + 2H_{2(g)}$

$2H_2O_{(g)}$



2 × H–H bonds are formed creating 2 × H_2 molecules.

Energy Change Calculations – Example #4:

The Electrolysis of Water to form Hydrogen and Oxygen

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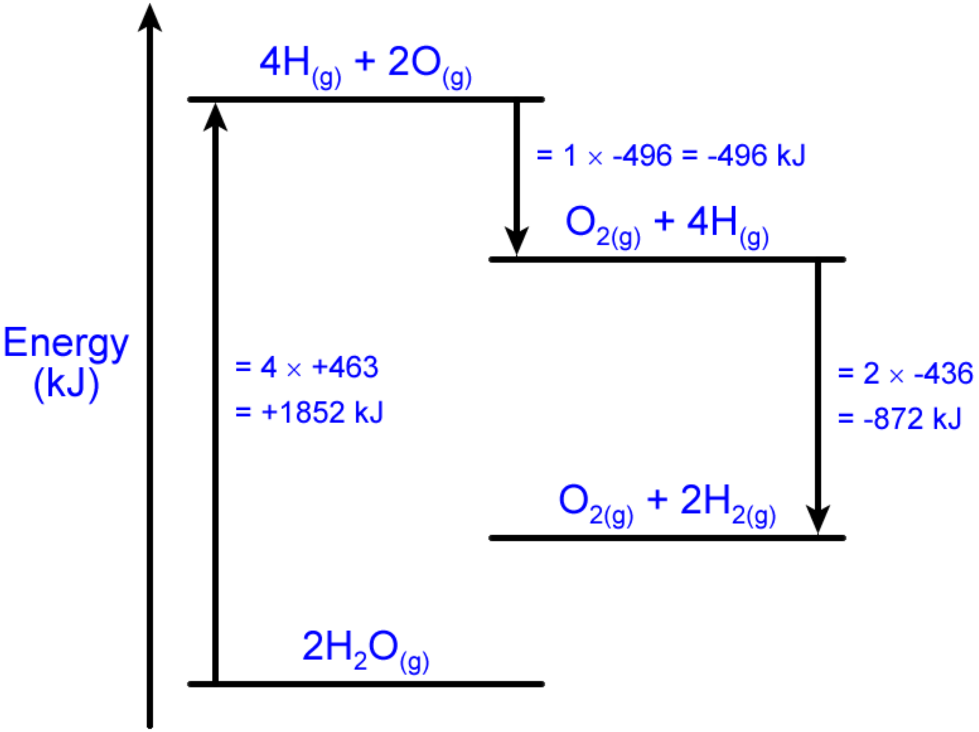


$4H_{(g)} + 2O_{(g)}$

= 1 × -496 = -496 kJ

$O_{2(g)} + 4H_{(g)}$

Energy (kJ)

= 4 × +463

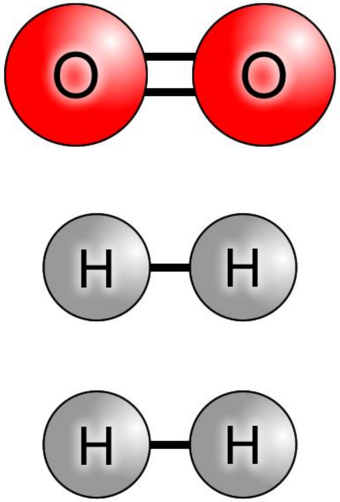
= +1852 kJ

= 2 × -436

= -872 kJ

$O_{2(g)} + 2H_{2(g)}$

$2H_2O_{(g)}$



Bond formation is *exothermic*: ΔH for this change is *negative*.

Energy Change Calculations – Example #4:

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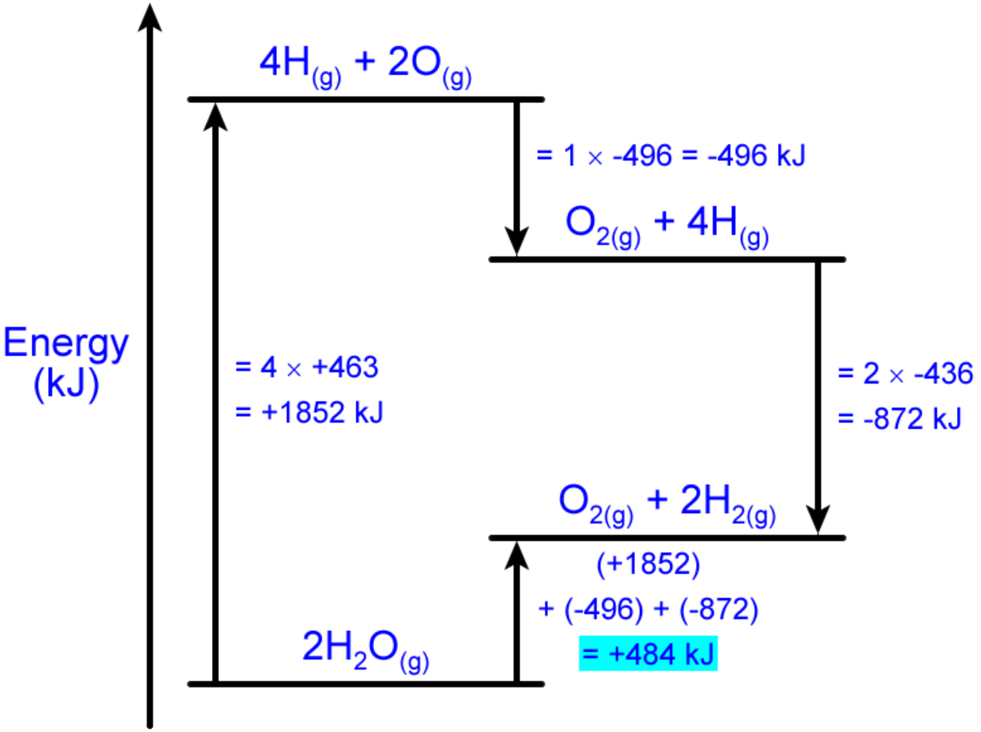

Energy (kJ)

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$O_{2(g)} + 4H_{(g)}$

= 4 × +463

= +1852 kJ

= 2 × -436

= -872 kJ

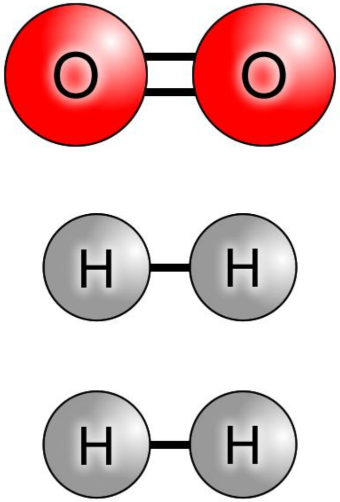
$O_{2(g)} + 2H_{2(g)}$

(+1852)

+ (-496) + (-872)

= +484 kJ

$2H_2O_{(g)}$



(+1852) + (−496) + (−872) = *+484 kJ*

Energy Change Calculations – Example #4:

The Electrolysis of Water to form Hydrogen and Oxygen

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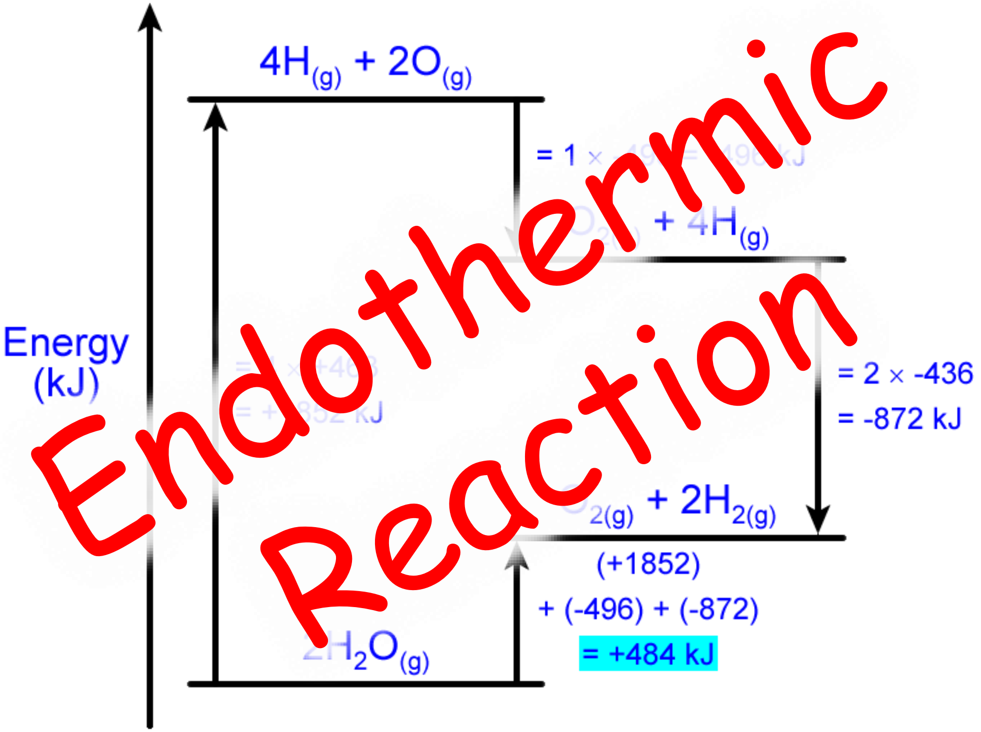
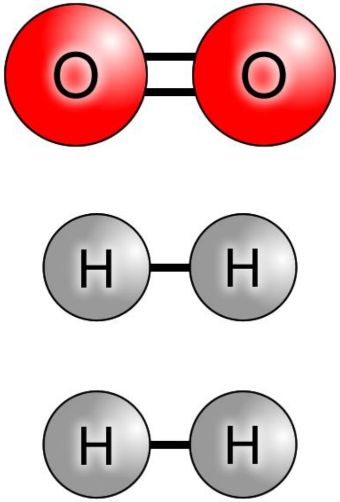


Endothermic Reaction



(+1852) + (−496) + (−872) = *+484 kJ*

Energy from Chemicals

Is a double covalent bond *twice as strong* as a single covalent bond?

Energy from Chemicals

- Study the bond energies given below:

C–C = 348 kJ   C=C = 612 kJ

O–O = 146 kJ   O=O = 496 kJ

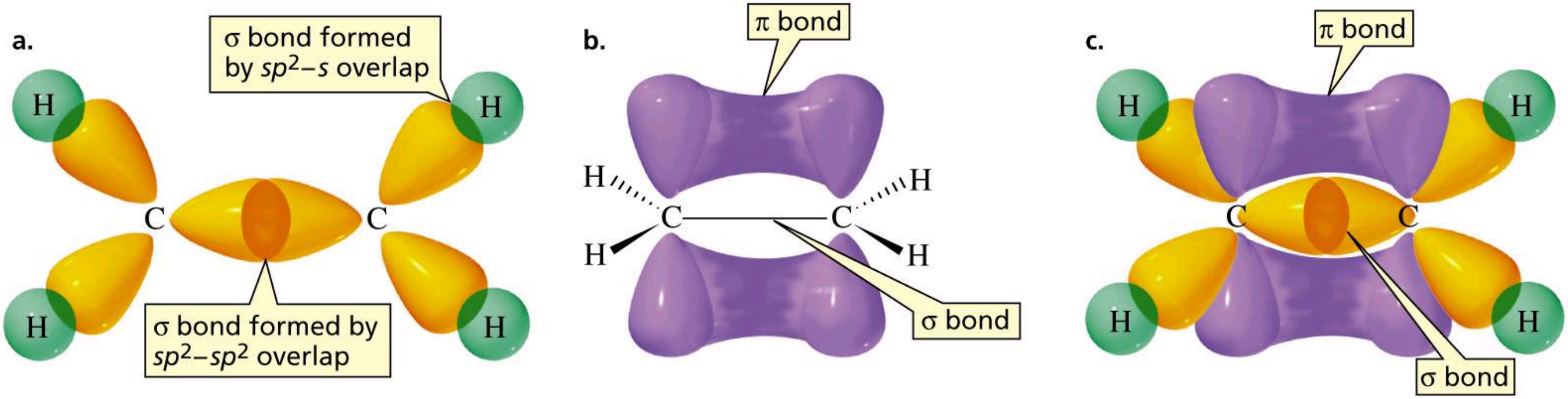
- Is a double covalent bond twice as strong as a single covalent bond?

- For carbon, 2 × C–C = 2 × 348 = 696 kJ
  696 kJ ≠ 612 kJ
  696 kJ > 612 kJ

- For oxygen, 2 × O–O = 2 × 146 = 292 kJ
  292 kJ ≠ 496 kJ
  292 kJ < 496 kJ

Energy from Chemicals

- Why is a double covalent bond not twice as strong as a single covalent bond?



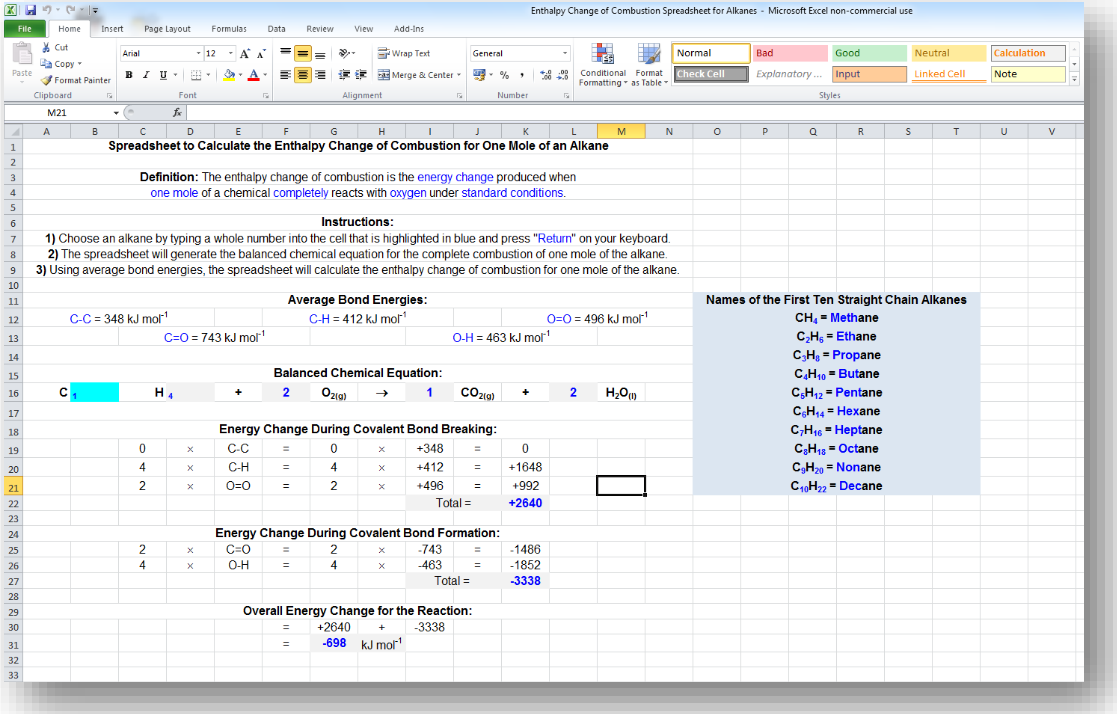
- Because there are different types of covalent bonds called σ-bonds (sigma bonds) and π-bonds (pi bonds).
- σ-bonds and π-bonds have different strengths.

Energy from Chemicals

The enthalpy change calculations seem to follow the same pattern. Can I automate the calculation in a *spreadsheet*?

Alkane Combustion Spreadsheet

- Alkanes are hydrocarbons with the general formula C_nH_{2n+2}.
- Alkanes react with oxygen to produce carbon dioxide and water.
- Design a spreadsheet that will, when the number of carbon atoms in the molecule is keyed in, automatically calculate the energy change for the combustion of one mole of an alkane.



Spreadsheet to Calculate the Enthalpy Change of Combustion for One Mole of an Alkane

Definition: The enthalpy change of combustion is the energy change produced when one mole of a chemical completely reacts with oxygen under standard conditions.

Instructions:

1) Choose an alkane by typing a whole number into the cell that is highlighted in blue and press "Return" on your keyboard.
2) The spreadsheet will generate the balanced chemical equation for the complete combustion of one mole of the alkane.
3) Using average bond energies, the spreadsheet will calculate the enthalpy change of combustion for one mole of the alkane.

Average Bond Energies:

C-C = 348 kJ mol^{-1}
C-H = 412 kJ mol^{-1}
O=O = 496 kJ mol^{-1}
C=O = 743 kJ mol^{-1}
O-H = 463 kJ mol^{-1}

Balanced Chemical Equation:

$C_1H_4 + 2\,O_{2(g)} \rightarrow 1\,CO_{2(g)} + 2\,H_2O_{(l)}$

Energy Change During Covalent Bond Breaking:

0 × C-C = 0 × +348 = 0
4 × C-H = 4 × +412 = +1648
2 × O=O = 2 × +496 = +992
Total = **+2640**

Energy Change During Covalent Bond Formation:

2 × C=O = 2 × -743 = -1486
4 × O-H = 4 × -463 = -1852
Total = **-3338**

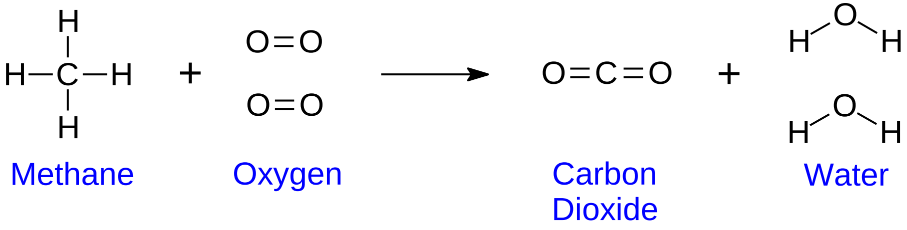
Overall Energy Change for the Reaction:

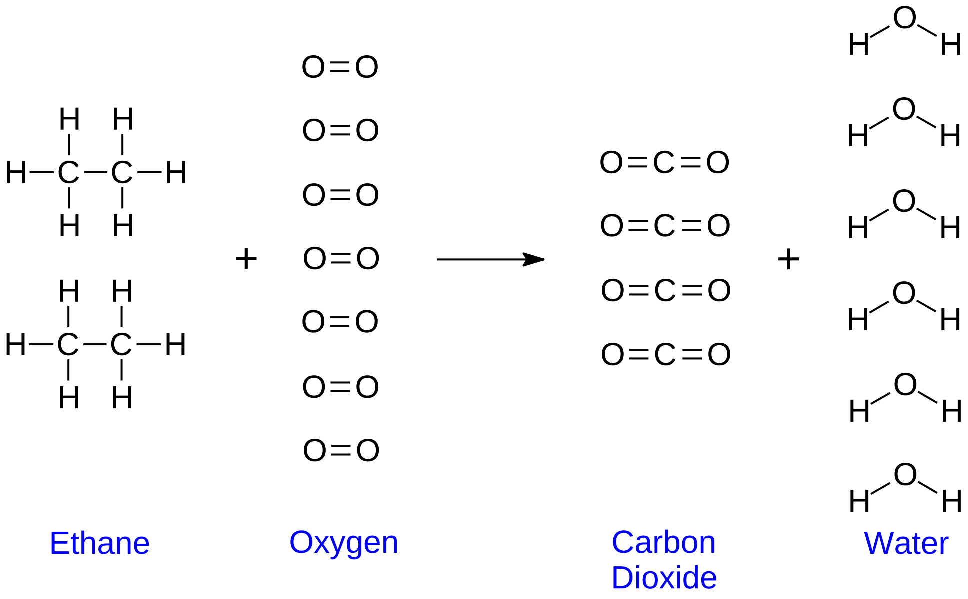
= +2640 + -3338
= **-698** kJ mol^{-1}

Names of the First Ten Straight Chain Alkanes

CH_4 = Methane
C_2H_6 = Ethane
C_3H_8 = Propane
C_4H_{10} = Butane
C_5H_{12} = Pentane
C_6H_{14} = Hexane
C_7H_{16} = Heptane
C_8H_{18} = Octane
C_9H_{20} = Nonane
$C_{10}H_{22}$ = Decane

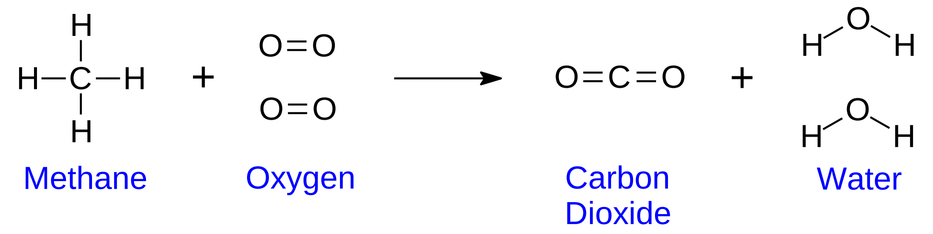
Alkane Combustion Spreadsheet

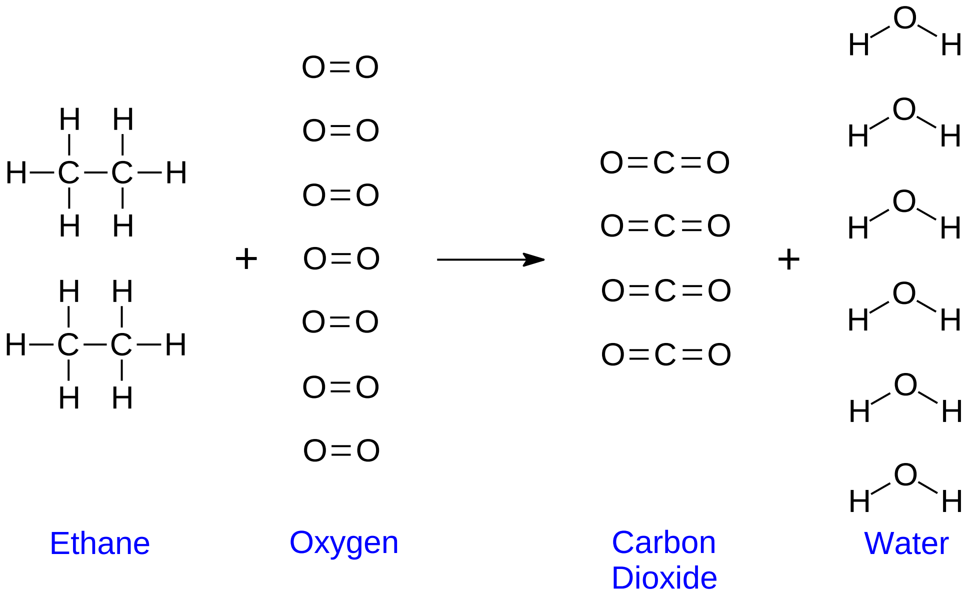




- Consider the combustion of methane and the combustion of ethane shown on the left.
- Let the number of carbon atoms = n.
- In relation to n, how many C–C bonds break?
- In relation to n, how many C–H bonds break?
- In relation to n, how many O=O bonds break?
- In relation to n, how many C=O bonds form?
- In relation to n, how many O–H bonds form?

Alkane Combustion Spreadsheet





- Consider the combustion of methane and the combustion of ethane shown on the left.
- Let the number of carbon atoms = n.
- In relation to n, how many C–C bonds break? $n - 1$
- In relation to n, how many C–H bonds break? $2n + 2$
- In relation to n, how many O=O bonds break? $1.5n + 0.5$
- In relation to n, how many C=O bonds form? $2n$
- In relation to n, how many O–H bonds form? $2n + 2$

Alkane Combustion Spreadsheet

- In summary, for an alkane with the general formula C_nH_{2n+2}:
  - The number of C–C bonds broken = n – 1.
  - The number of C–H bonds broken = 2n + 2.
  - The number of O=O bonds broken = 1.5n + 0.5.
  - The number of C=O bonds formed = 2n.
  - The number of O–H bonds formed = 2n + 2.
  - Average C–C bond energy = 348 kJ / mol.
  - Average C–H bond energy = 411 kJ / mol.
  - Average O=O bond energy = 498 kJ / mol.
  - Average C=O bond energy = 804 kJ / mol.
  - Average O–H bond energy = 463 kJ / mol.
- Remember, bond breaking is endothermic, ΔH is positive, while bond formation is exothermic, Δ is negative.

Alkane Combustion Spreadsheet

- How do your results compare to the experimental data?

Name of Straight Alkane	Formula of Straight Chain Alkane	Physical State at Room Temperature and Pressure	Standard Molar Enthalpy Change of Combustion at 298 K / kJ/mol
Methane	CH_4	Gas	–890
Ethane	C_2H_6	Gas	–1560
Propane	C_3H_8	Gas	–2220
Butane	C_4H_{10}	Gas	–2877
Pentane	C_5H_{12}	Liquid	–3509
Hexane	C_6H_{14}	Liquid	–4163

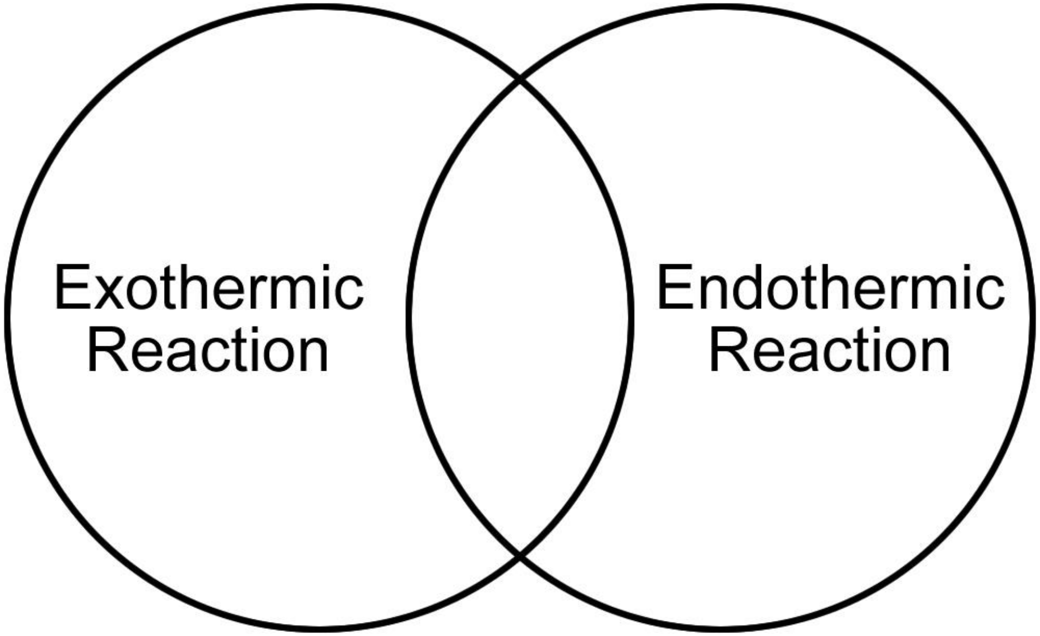
- Identify any possible sources of error in the calculations.

Energy from Chemicals

Compare and contrast *exothermic* and *endothermic* reactions. In what ways are they *similar*? In what ways are they *different*?

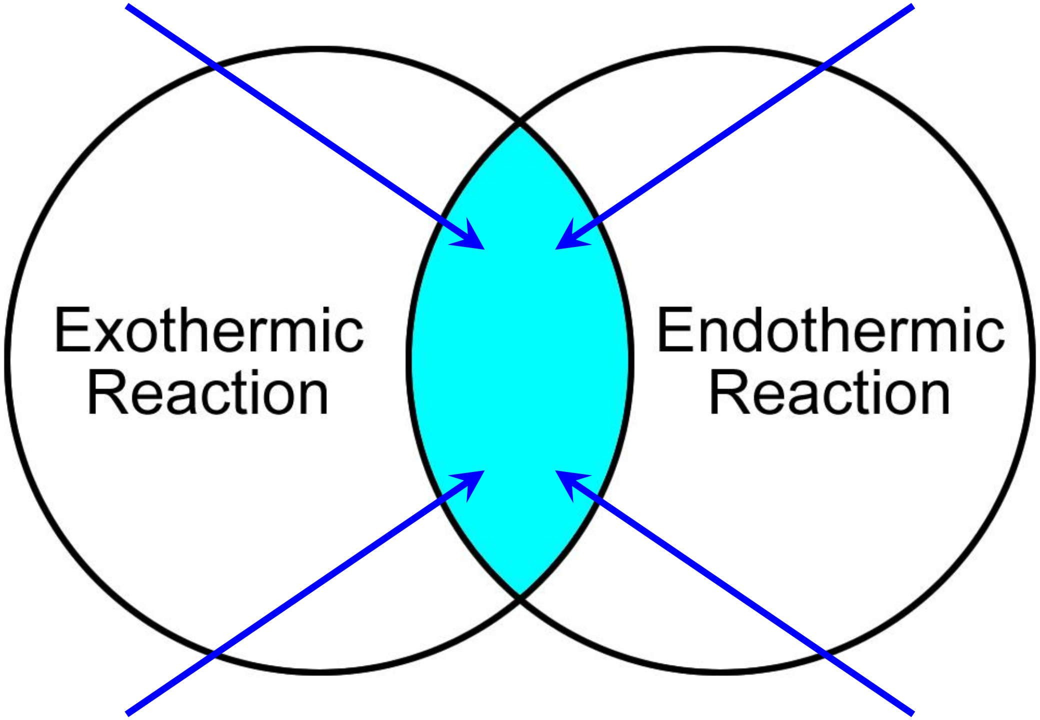
Higher Order Thinking Skills

Compare and Contrast Exothermic and Endothermic Reactions



Higher Order Thinking Skills

Compare and Contrast Exothermic and Endothermic Reactions



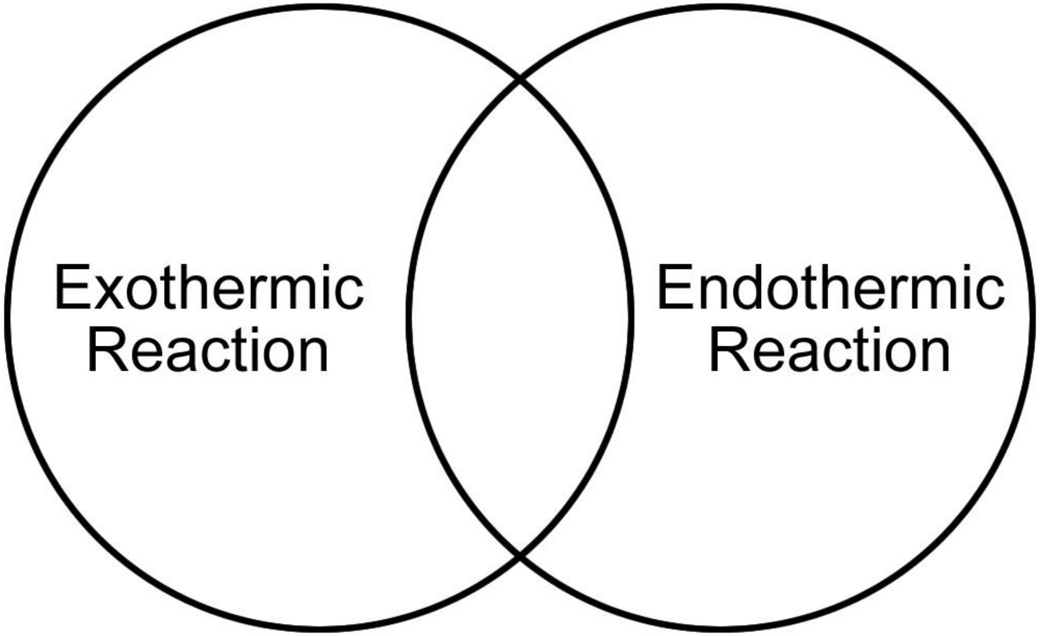
Exothermic Reaction

Endothermic Reaction

- Reaction involves bond formation.
- Reaction involves bond breaking.
- Reaction has an activation energy.
- A catalyst will lower the activation energy.

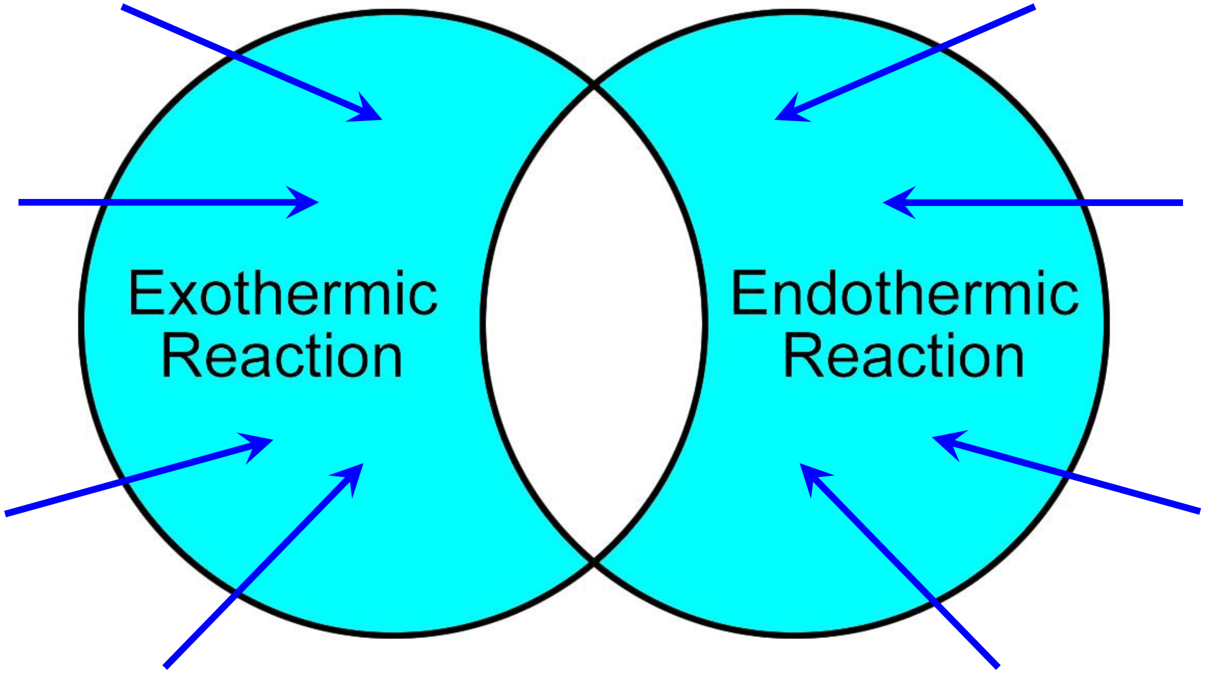
Higher Order Thinking Skills

Compare and Contrast Exothermic and Endothermic Reactions



Higher Order Thinking Skills

Compare and Contrast Exothermic and Endothermic Reactions



Exothermic Reaction

- Energy absorbed to break bonds is less than the energy released when bonds form.
- Energy of reactants is greater than energy of products.
- ΔH is negative.
- Energy released into the surroundings (temperature increases).

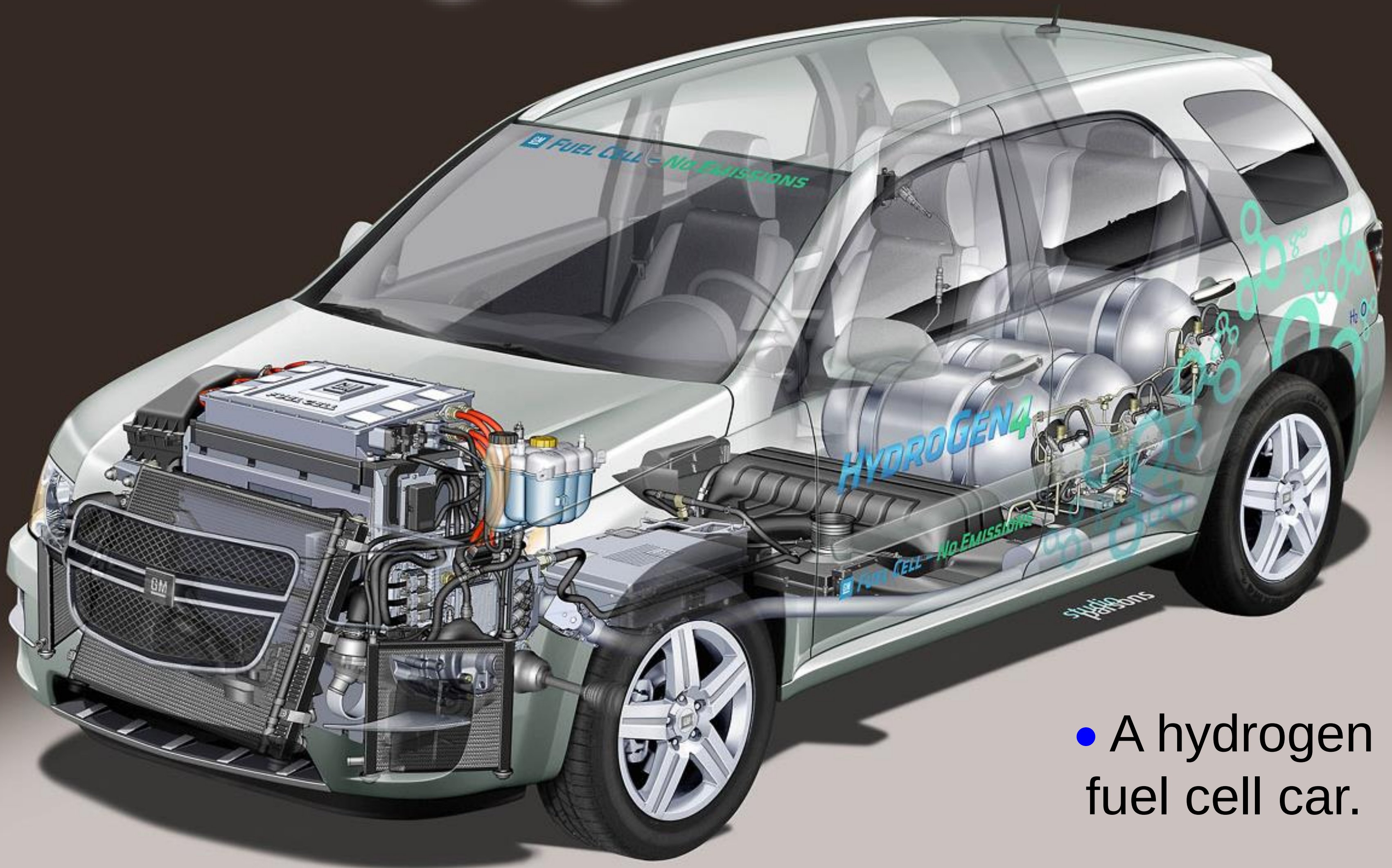
Endothermic Reaction

- Energy absorbed to break bonds is greater than the energy released when bonds form.
- Energy of reactants is less than energy of products.
- ΔH is positive.
- Energy absorbed from the surroundings (temperature decreases).

Energy from Chemicals

Hydrogen fuel cells generate electricity without causing any pollution. How do they work?

Hydrogen Fuel Cells

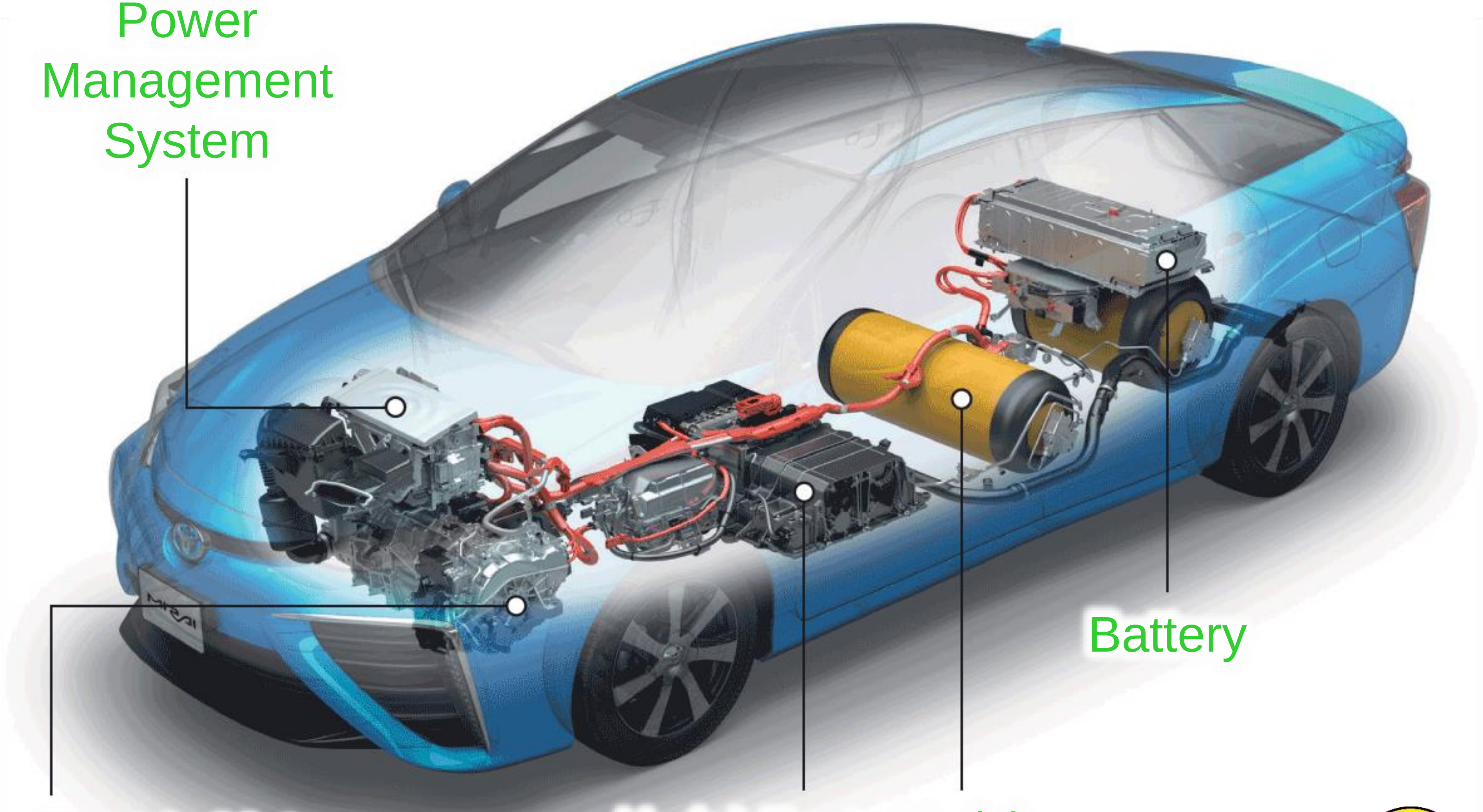


- A hydrogen fuel cell car.

Hydrogen Fuel Cells

- The Toyota Mirai – the world’s first commercial hydrogen fuel cell car – went on sale in Japan on 15th December 2014.

Hydrogen Fuel Cells



Power Management System

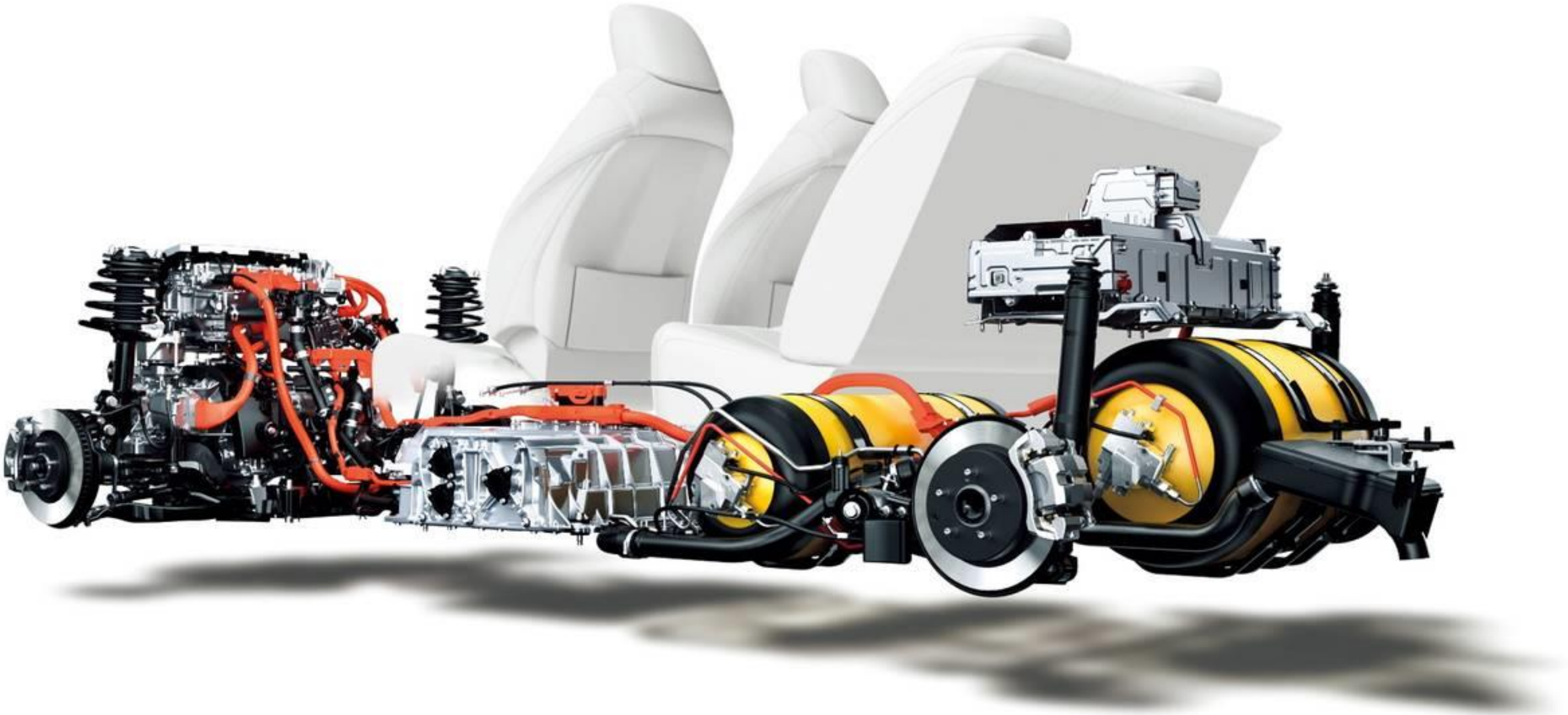
Battery

Electric Motor

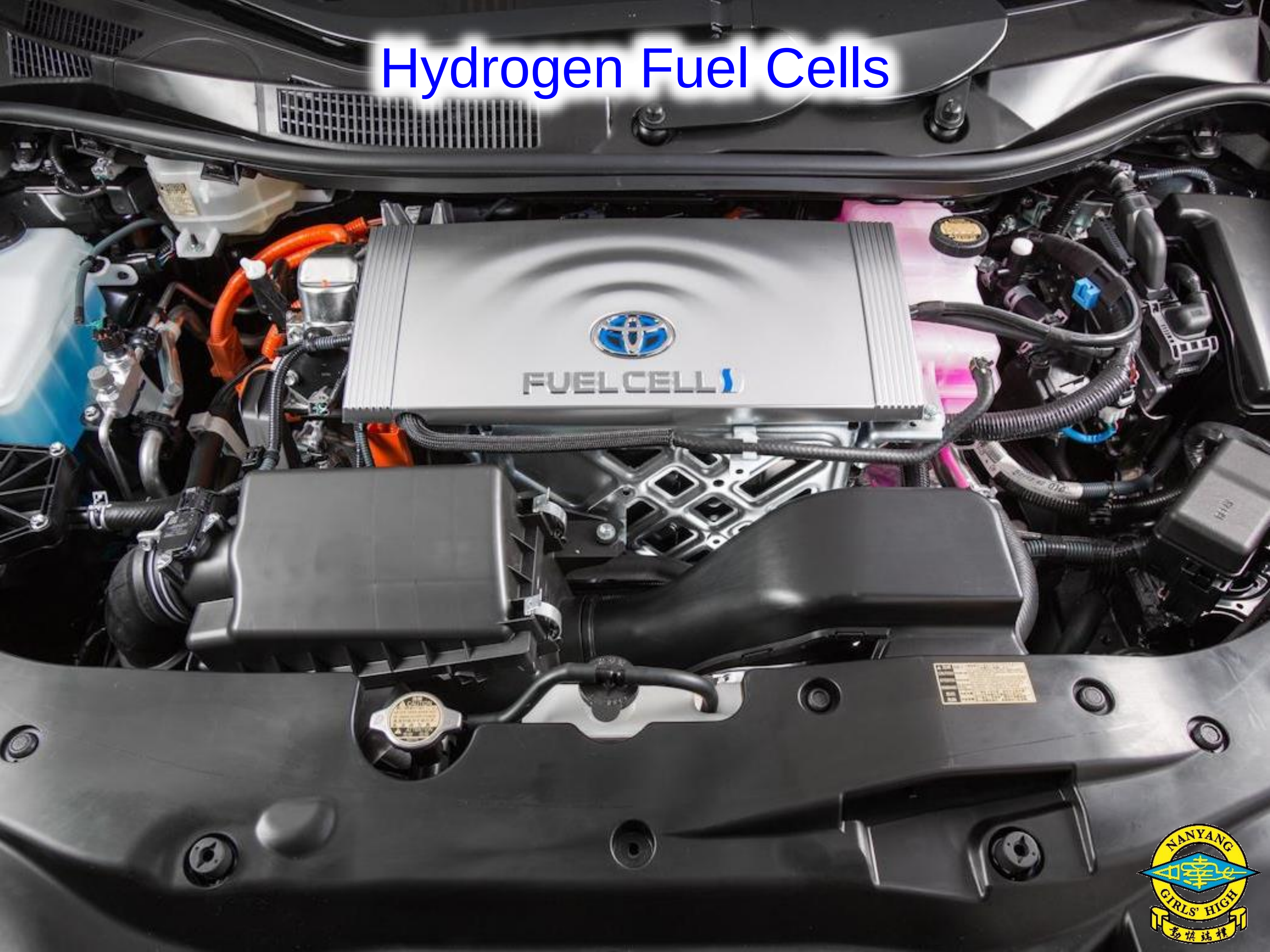
H_2 (g) Fuel Cell

H_2 (g) Tank

Hydrogen Fuel Cells



Hydrogen Fuel Cells



Hydrogen Fuel Cells

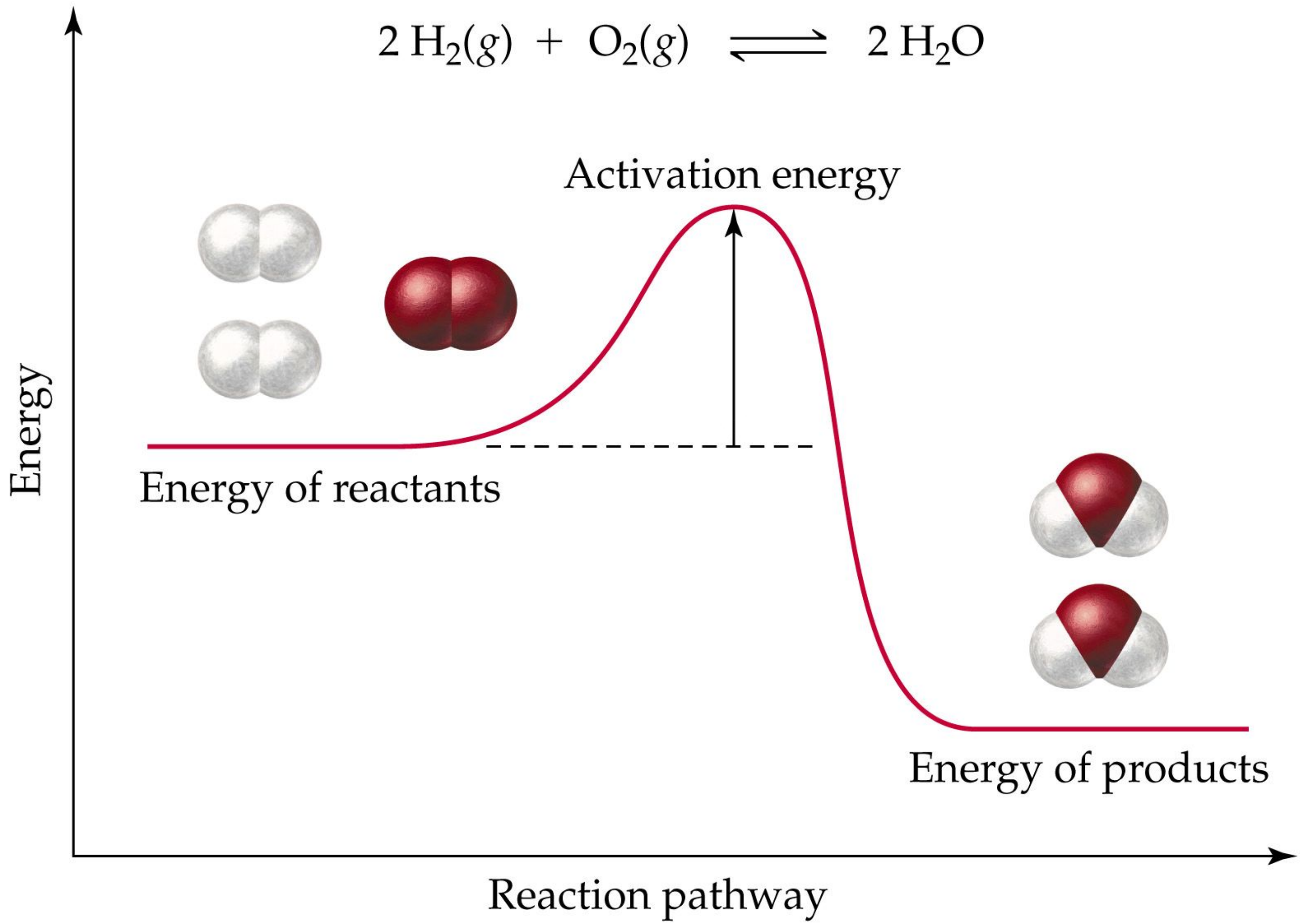
Hydrogen Fuel Cells

- When hydrogen reacts with oxygen, an exothermic reaction takes place that produces water as the only reaction product.

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(l)} \quad \Delta H = -572 \text{ kJ}$$

- In a hydrogen fuel cell, the hydrogen reacts with oxygen to produce water and energy in the form of electricity. The electrical energy can be used to do useful work, such as driving an electric motor to move a car.

Hydrogen Fuel Cells



Hydrogen Fuel Cells

- In the hydrogen fuel cell, a catalyst causes molecular hydrogen to break-down into hydrogen ions and electrons.

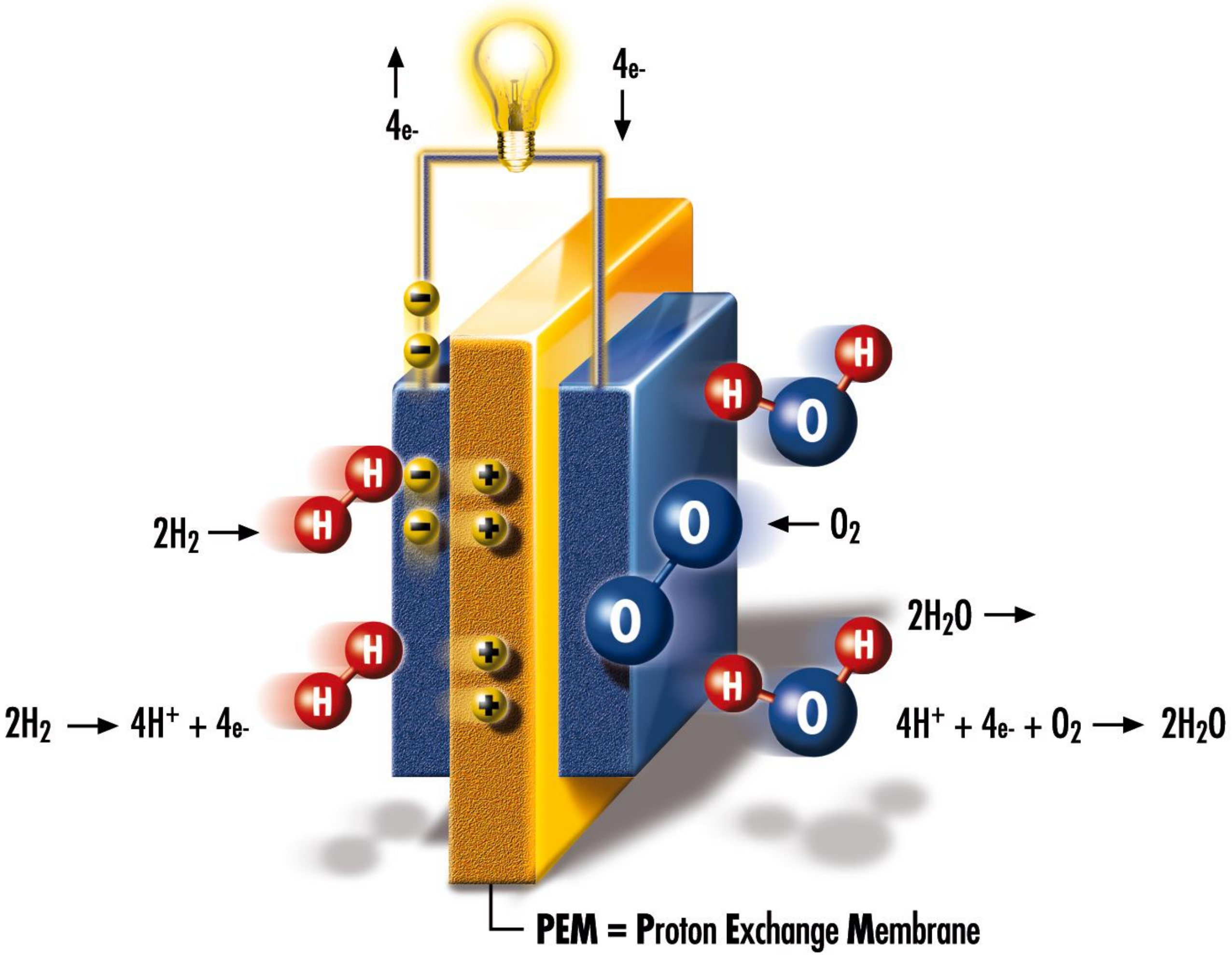
$$H_{2(g)} \rightarrow 2H^{+}_{(aq)} + 2e^{-}$$

- The electrons pass through metal wires in an external circuit where they are made to do useful work, such as drive an electric motor. The hydrogen ions pass through a special membrane.

- On the other side of the membrane, the hydrogen ions, electrons and oxygen combine together to form water.

$$4H^{+}_{(aq)} + 4e^{-} + O_{2(g)} \rightarrow 2H_2O_{(l)}$$

Hydrogen Fuel Cells



Hydrogen Fuel Cells

- Hydrogen required for the hydrogen fuel cell is obtained by *cracking* long-chain hydrocarbons from crude oil, *e.g.*

$$C_{20}H_{42(s)} \rightarrow C_8H_{16(l)} + C_{12}H_{24(l)} + H_{2(g)}$$

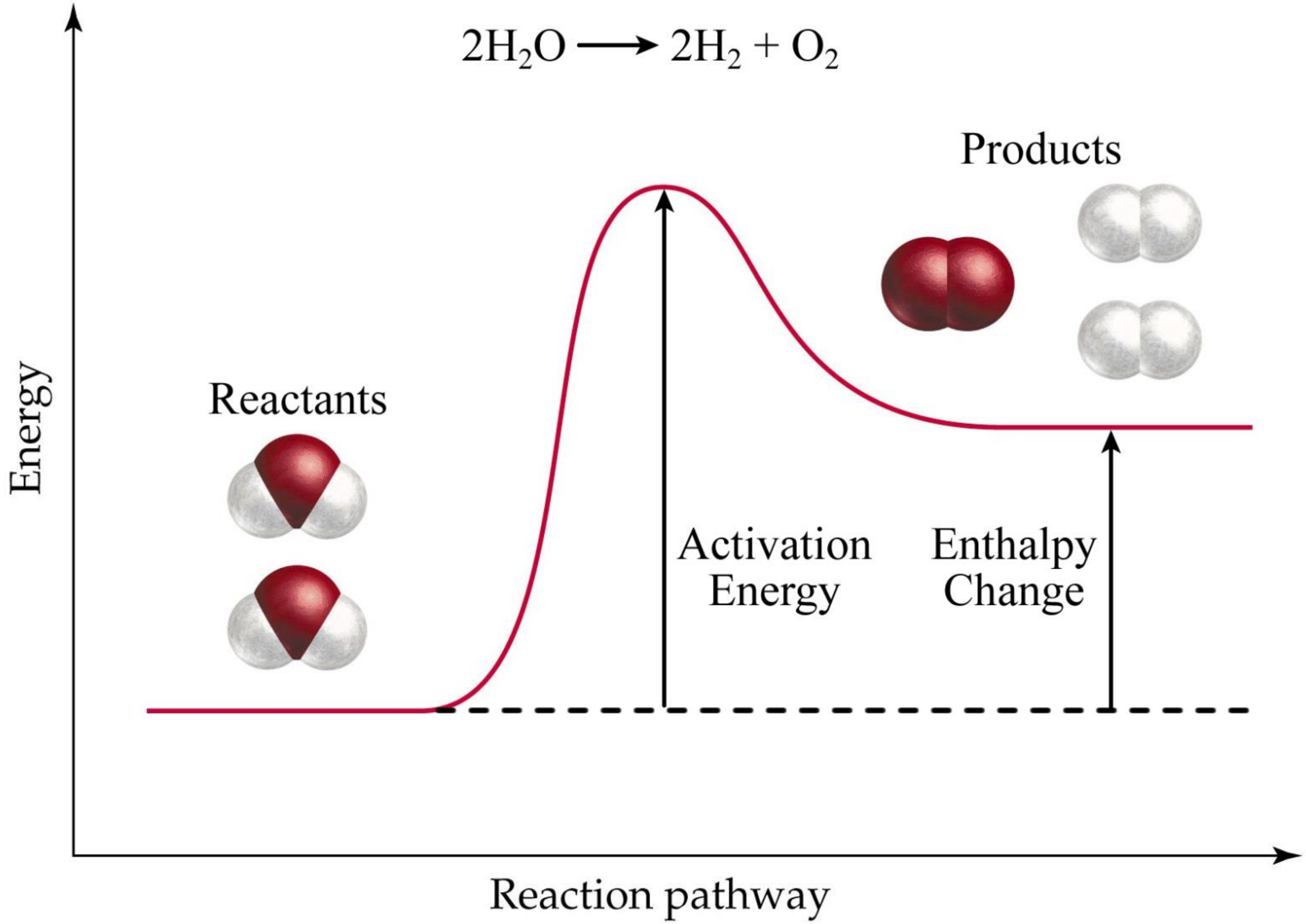
- Hydrogen can also be obtained from the *electrolysis* of water (the decomposition of water by electricity):

$$2H_2O_{(l)} \rightarrow 2H_{2(g)} + O_{2(g)}$$

- Remember, the hydrogen that is produced by cracking long-chain hydrocarbons is also used in the manufacture of ammonia – NH_3.

- The hydrogen must be stored with the hydrogen fuel cell. The large scale storage of hydrogen is hazardous because it is a highly flammable gas.

Hydrogen Fuel Cells



- The electrolysis of water is an *endothermic* process.

Hydrogen Fuel Cells

- Oxygen required for the hydrogen fuel cell is taken directly from the Earth's atmosphere, which is 21% oxygen.

Energy from Chemicals

Why is the hydrogen fuel cell considered to be a *clean* source of energy?

Energy from Chemicals

Why is the hydrogen fuel cell considered to be a *clean* source of energy?

- The only chemical product of the reaction is *water*. Water is not considered to be a pollutant. Compare this to the chemicals that are produce by the internal combustion engine.

Energy from Chemicals

Is the hydrogen fuel cell really such a *clean* source of energy?

Energy from Chemicals

Is the hydrogen fuel cell really such a *clean* source of energy?

- Consider how the raw materials (hydrogen and oxygen) for the fuel cell are obtained. To what extent is atmospheric pollution a possible consequence?

Energy from Chemicals

How would I *design an experiment* to measure the energy change of a reaction?

Experimental Design

- The combustion of an alcohol is an exothermic reaction. For example:

methanol $2CH_3OH_{(l)} + 3O_{2(g)} \rightarrow 2CO_{2(g)} + 4H_2O_{(l)}$

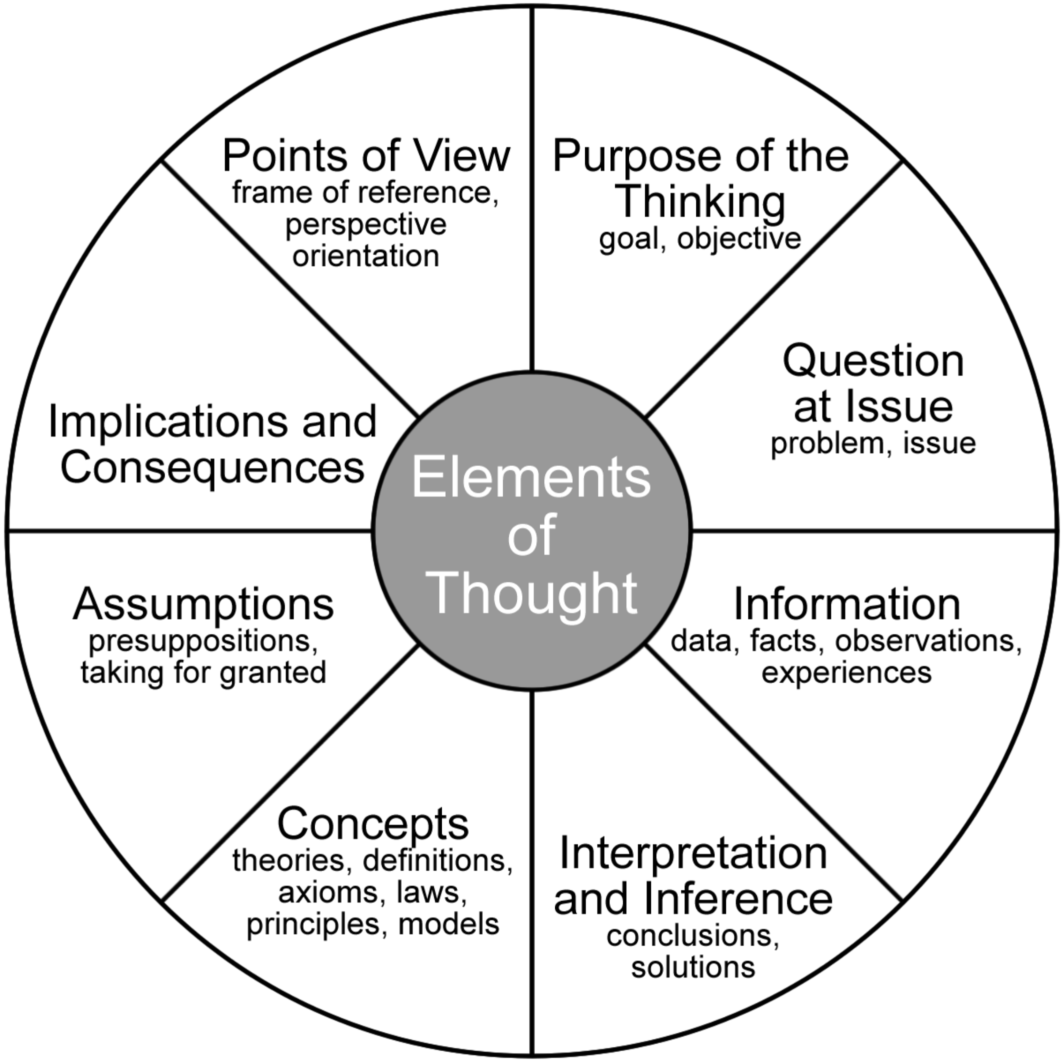
ethanol $C_2H_5OH + 3O_{2(g)} \rightarrow 2CO_{2(g)} + 3H_2O_{(l)}$

propan-1-ol $2C_3H_7OH + 9O_{2(g)} \rightarrow 6CO_{2(g)} + 8H_2O_{(l)}$

- Design an experiment to determine which alcohol, methanol, ethanol or propan-1-ol, releases the greatest amount of energy when one gram of the alcohol is completely burned in air.

Experimental Design

Paul’s Wheel of Reason



Elements of Thought

Purpose of the Thinking
goal, objective

Question at Issue
problem, issue

Information
data, facts, observations, experiences

Interpretation and Inference
conclusions, solutions

Concepts
theories, definitions, axioms, laws, principles, models

Assumptions
presuppositions, taking for granted

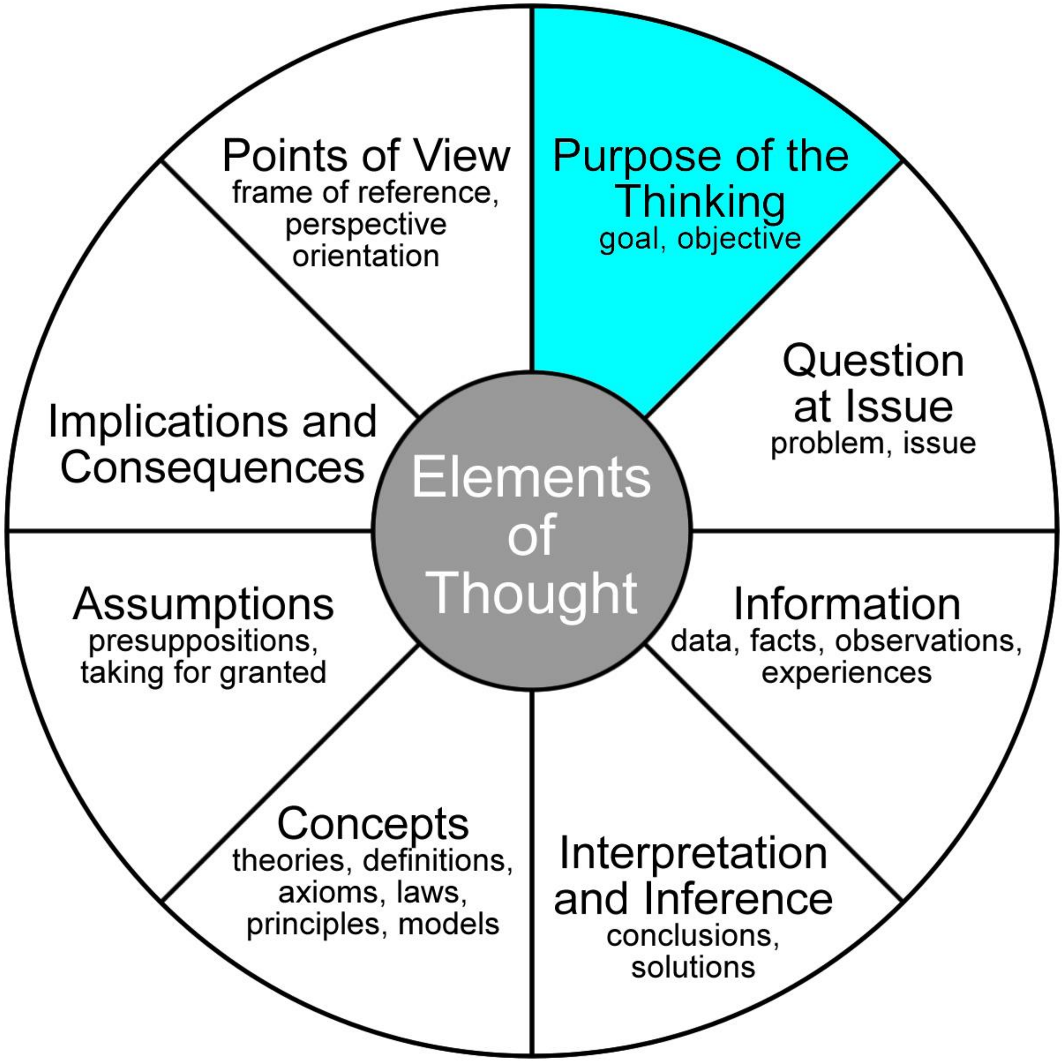
Implications and Consequences

Points of View
frame of reference, perspective orientation

- Experiments should be designed using critical thinking skills.

Experimental Design

Paul’s Wheel of Reason



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Points of View
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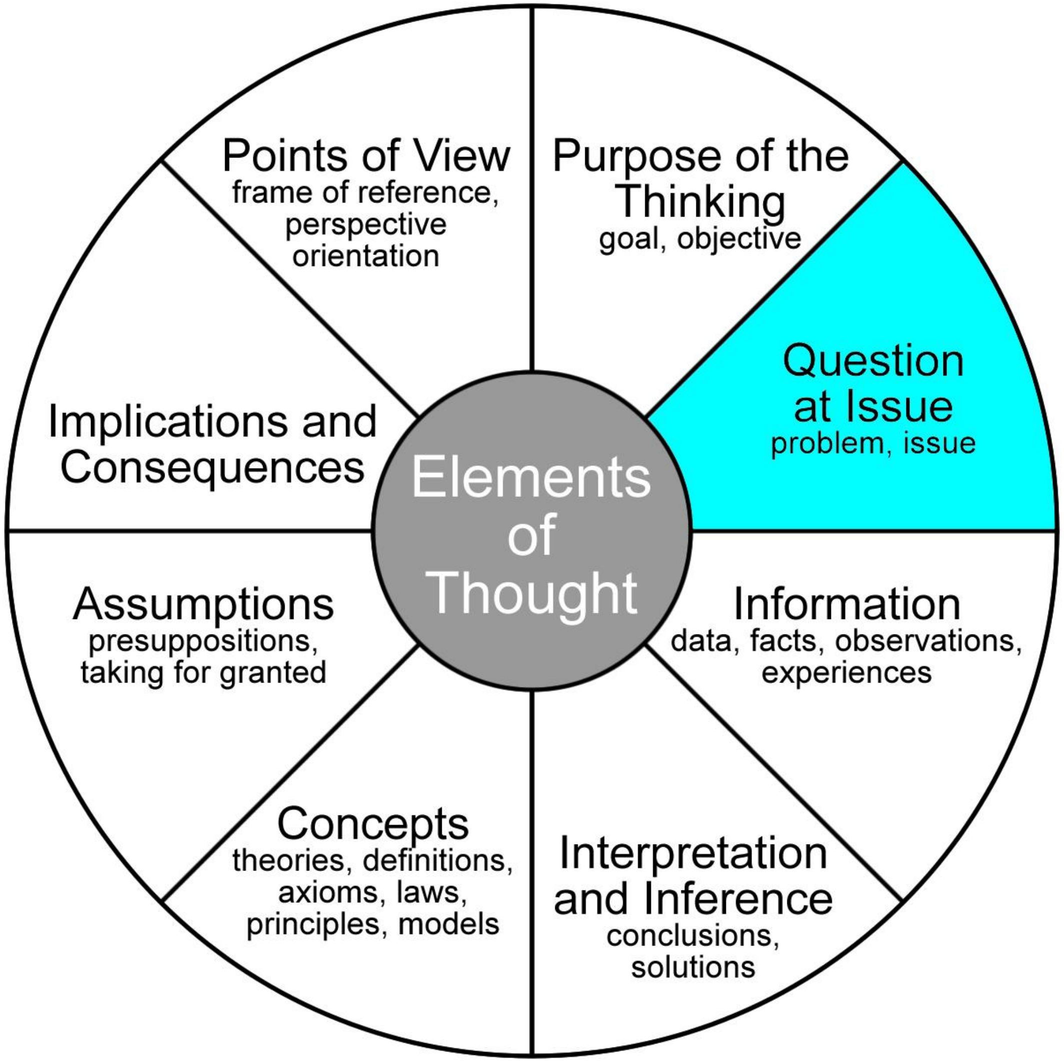
- What is the reason for performing the experiment?
- What hypothesis is being proved / disproved by performing the experiment?

Experimental Design

Paul's Wheel of Reason



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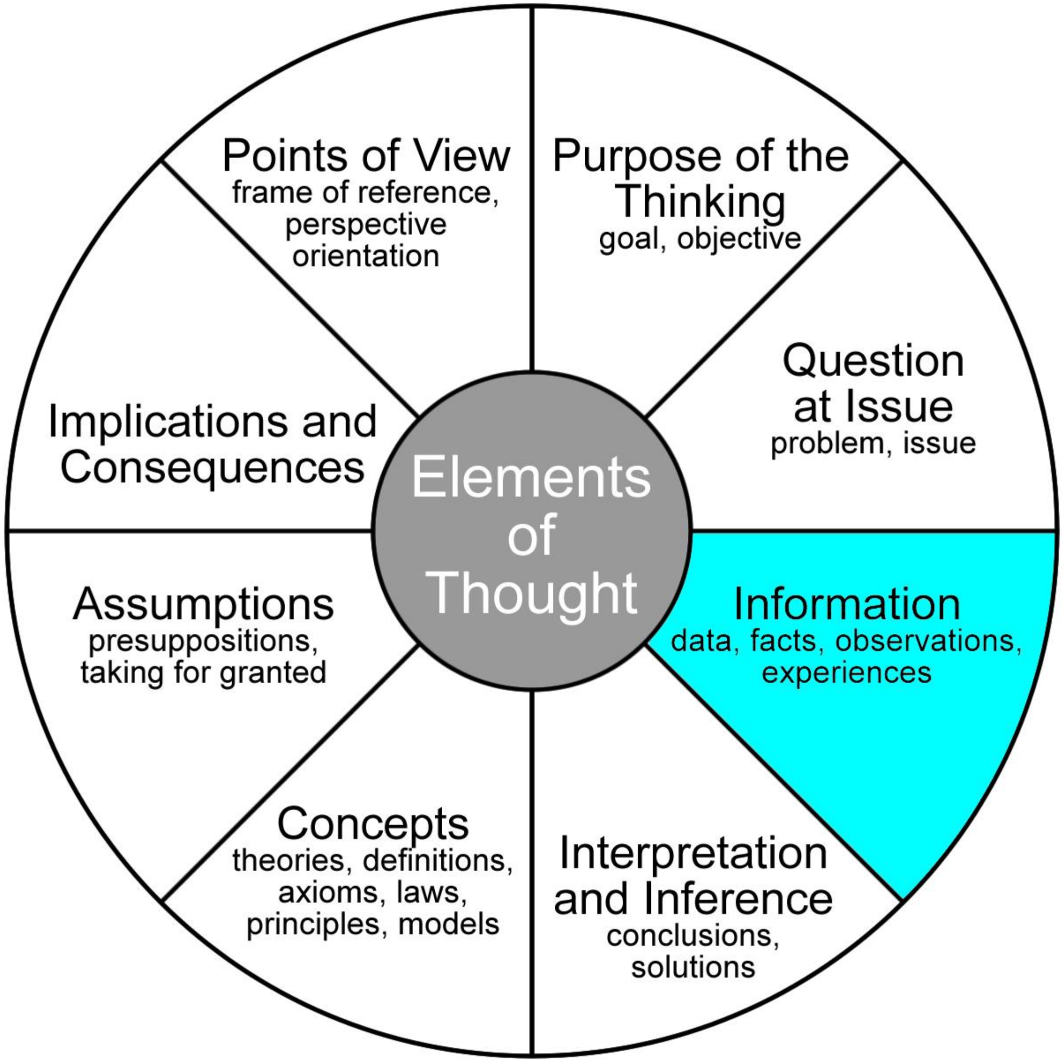
Implications and Consequences

- Clearly state exactly what the experiment is being designed to do.
- What is the problem statement for the experiment?

NANYANG GIRLS' HIGH

Experimental Design

Paul’s Wheel of Reason



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Implications and Consequences

Points of View
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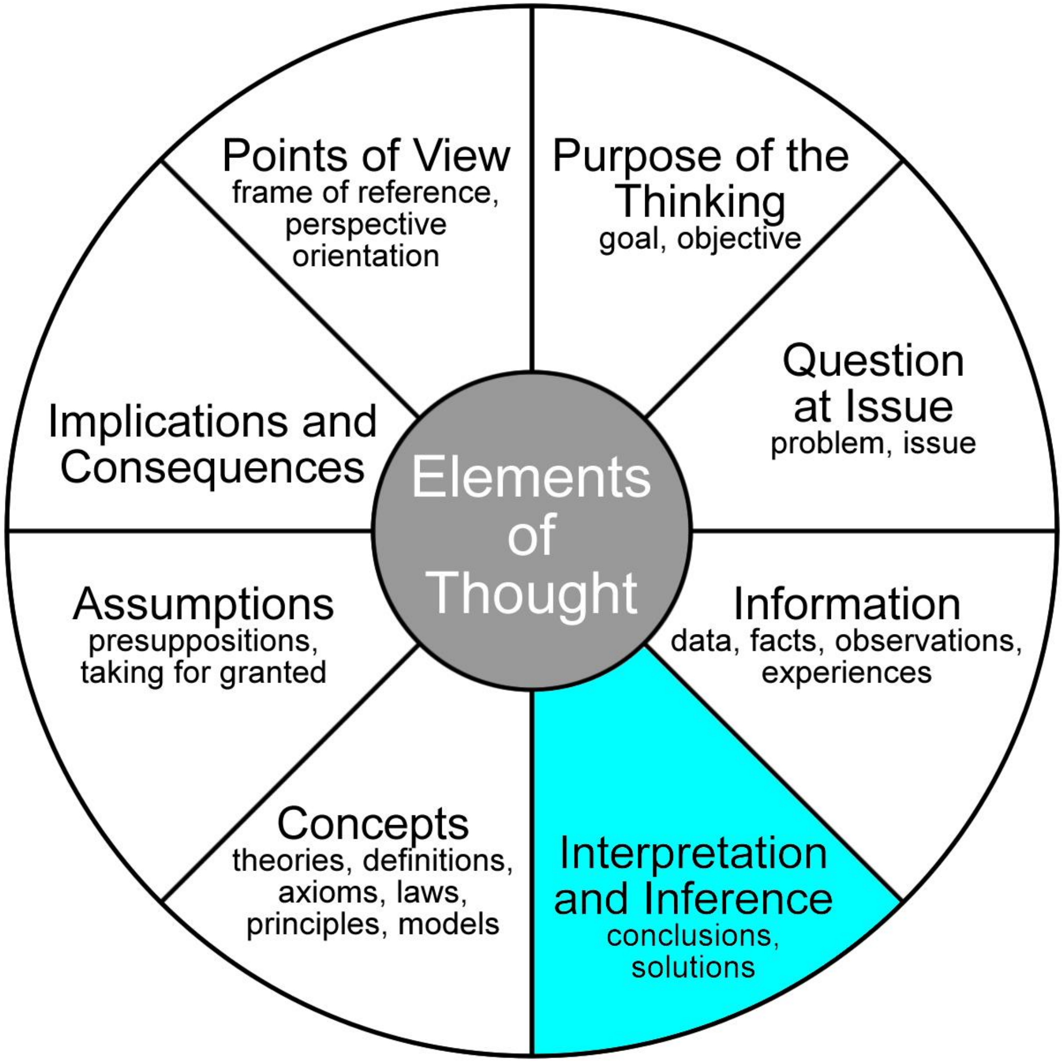
- What essential information needs to be collected by the experiment?
- What observations need to be made?
- What are the different variables (independent, dependent and control)?

Experimental Design

Paul's Wheel of Reason



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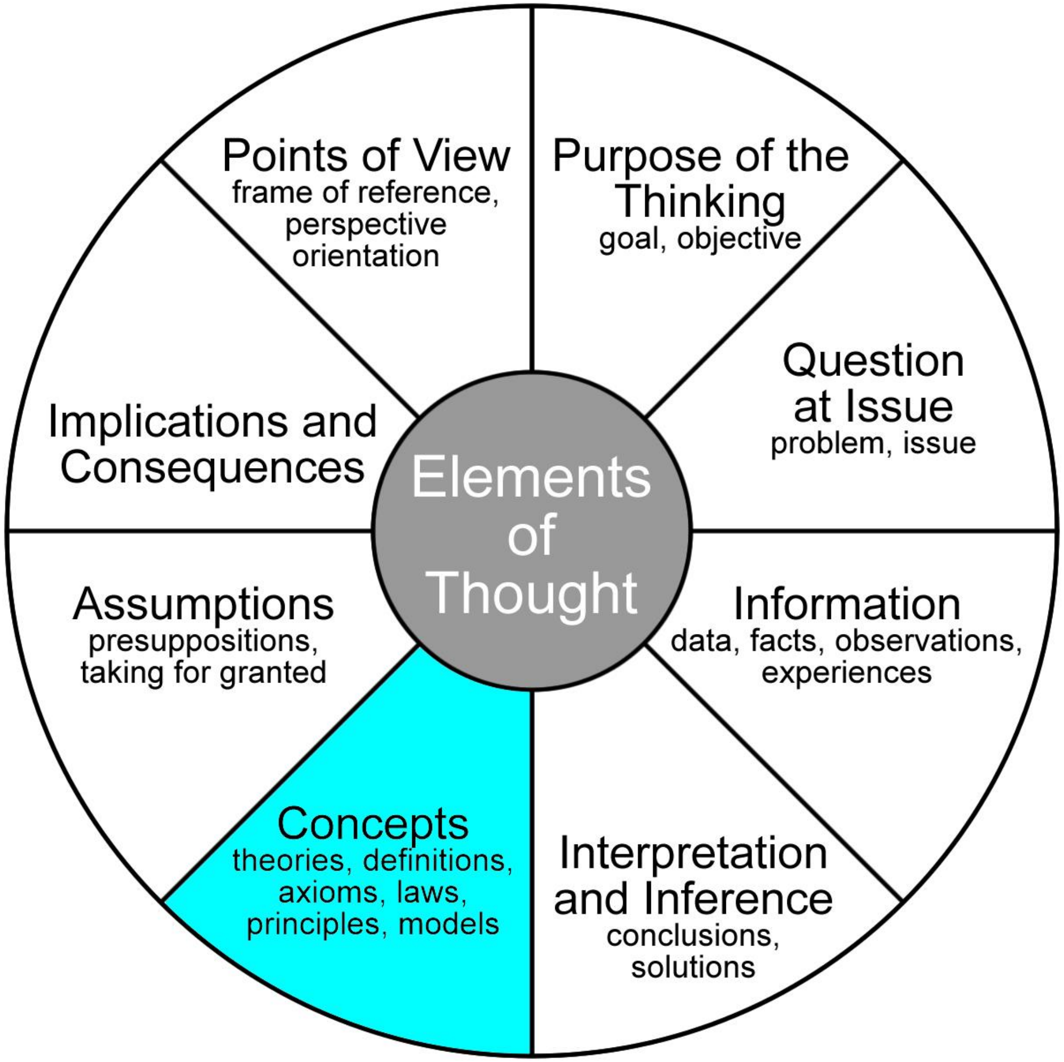
- How must the data from the experiment be manipulated and presented?
- What calculations (if any) need to be done?
- What graphs (if any) need to be plotted?

Experimental Design

Paul’s Wheel of Reason



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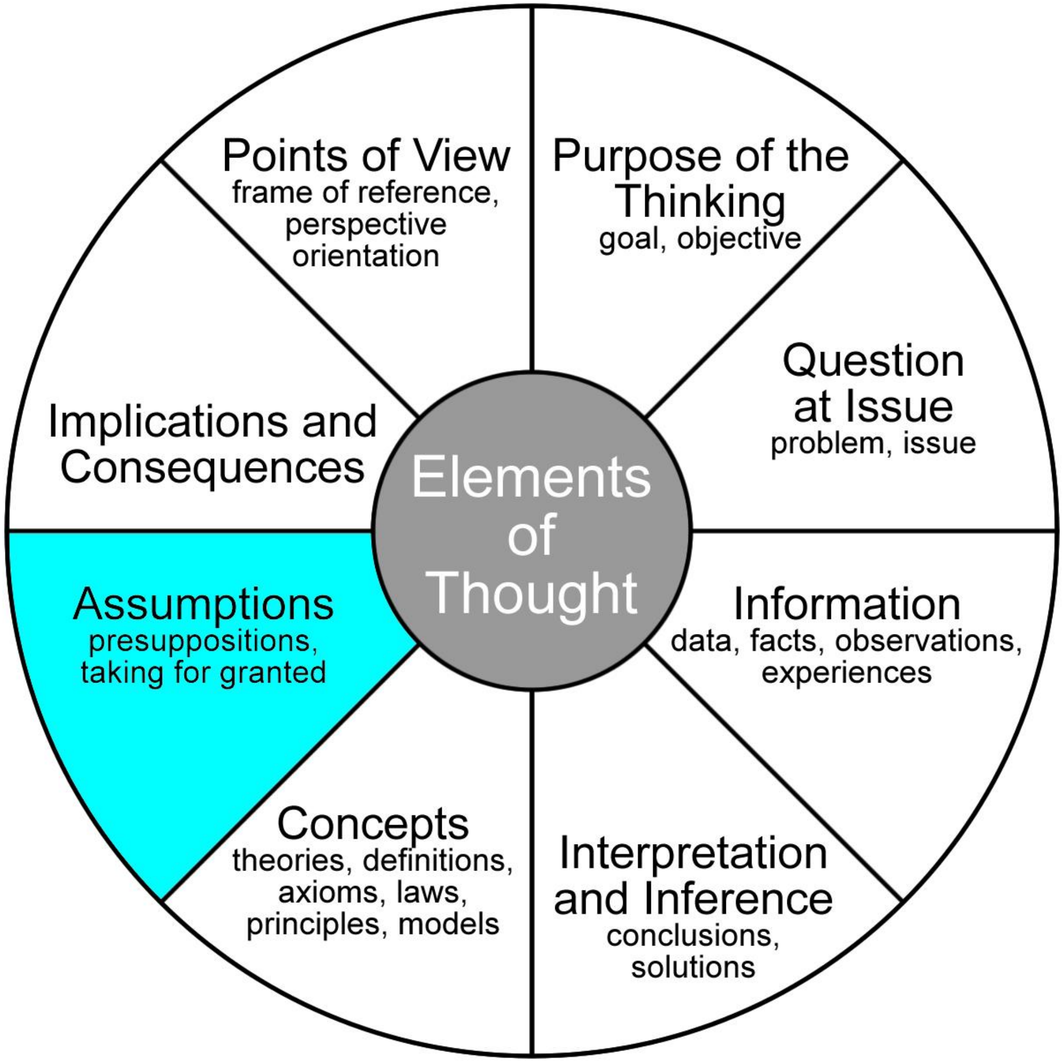
- What essential concept / law does the design of the experiment rely on? For example, the concept of energy and the law that energy can not be created or destroyed, but only converted from one form into another.

Experimental Design

Paul’s Wheel of Reason



- What assumptions are being made when designing the experiment?
- Is it reasonable to make these assumptions?
- Identifying assumptions can help to identify errors.

Experimental Design

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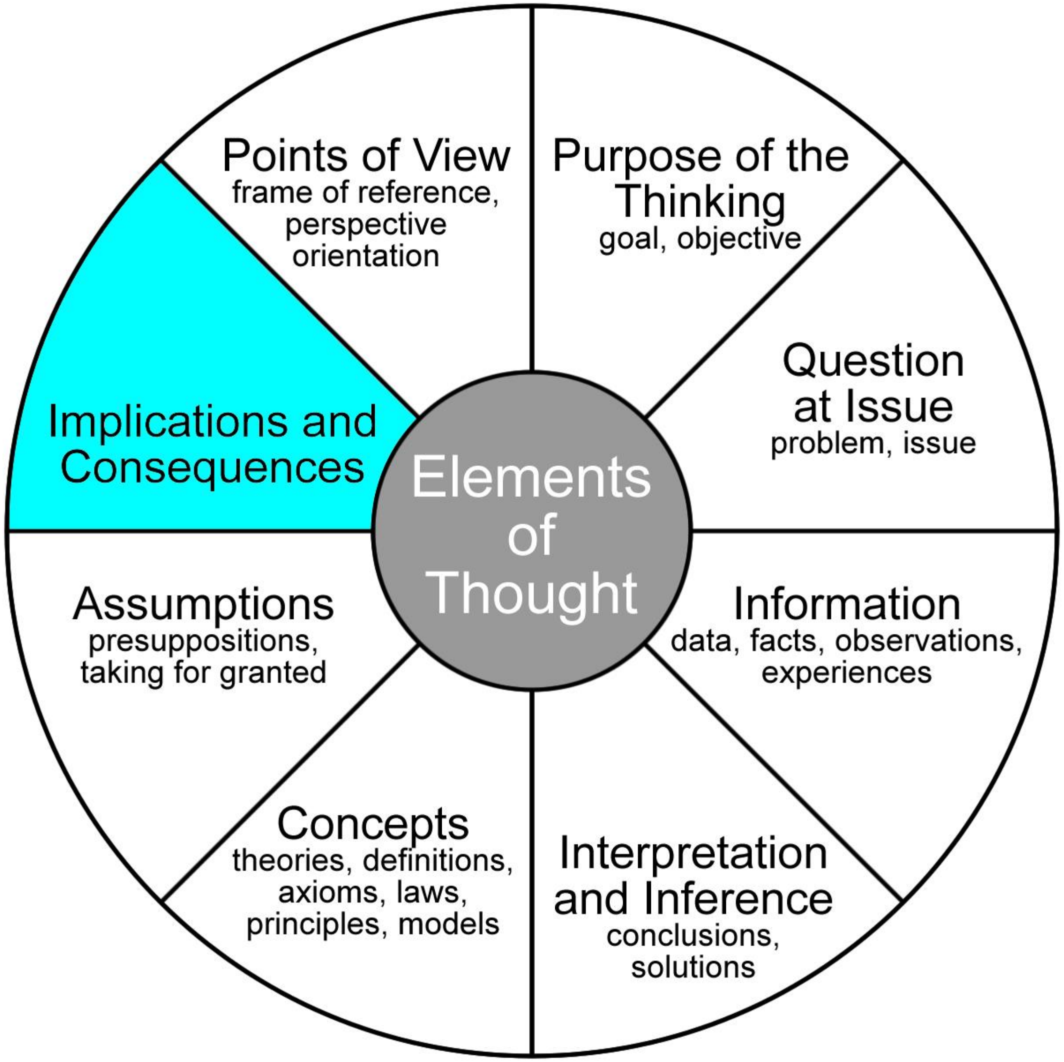

Points of View
frame of reference, perspective orientation

Purpose of the Thinking
goal, objective

Question at Issue
problem, issue

Information
data, facts, observations, experiences

Interpretation and Inference
conclusions, solutions

Concepts
theories, definitions, axioms, laws, principles, models

Assumptions
presuppositions, taking for granted

Implications and Consequences

Elements of Thought

- What are the implications and consequences if the results obtained from the experiment prove / disprove a specific hypothesis?

Experimental Design

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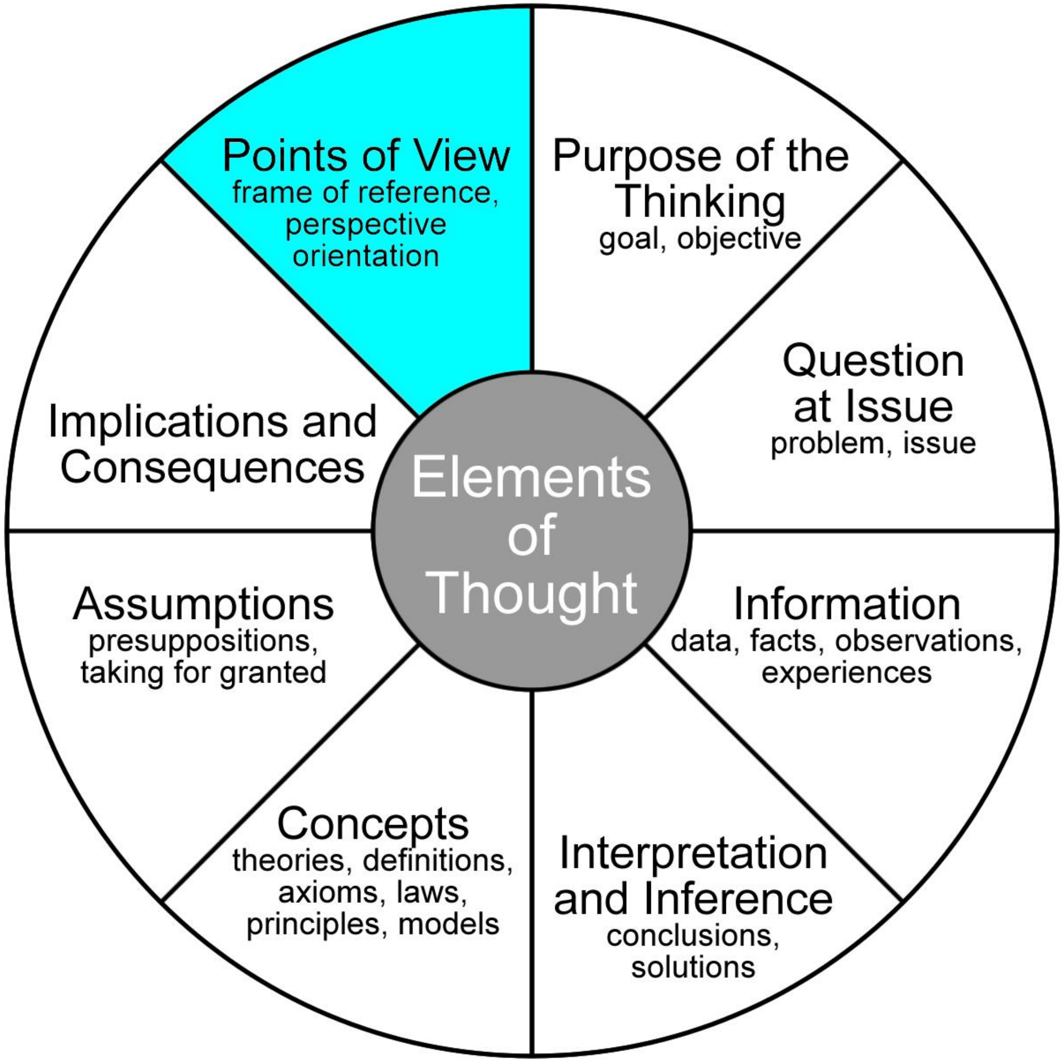

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Implications and Consequences

- Is there more than just one way of doing this experiment?
- If there is more than one way, which way is the most valid and reliable?

Experimental Design

- What is your hypothesis or problem statement?
- What are the variables? Which variables must be kept constant? What is your independent (input) variable? What is your dependent (output) variable?
- What is the big idea? Provide a general overview of the method.
- What apparatus is required to perform the experiment?
- How is the apparatus set-up? Provide a clear, labelled diagram.
- How will you perform the experiment? Provide step-by-step instructions, including how variables will be controlled and what measurements will be taken.
- How will data from the experiment will be recorded? Provide an example results table that includes clear headings and units.
- How will the experimental data be manipulation? What graphs need to be plotted? What calculations need to be performed?
- What errors are associated with the experiment? What effect do these errors have on your results?

Experimental Design

Problem Statement

- Does the alcohol containing the greater number of carbon atoms per molecule release more energy when one gram of it is burned?

Variables

- Mass of water in the copper calorimeter (constant).
- Mass of alcohol burned during the experiment (independent variable).
- Temperature change of the water in the copper calorimeter (dependent variable).

Experimental Design

General Overview of the Method

A known mass of alcohol is burned and used to heat a known mass of water, causing its temperature to increase by a known amount. Using the equation $E = m \times c \times \Delta T$, the heat energy absorbed by the water is calculated. It is assumed that the energy used to heat the water originated from combustion of the alcohol, therefore the energy released per unit mass of the alcohol can be calculated.

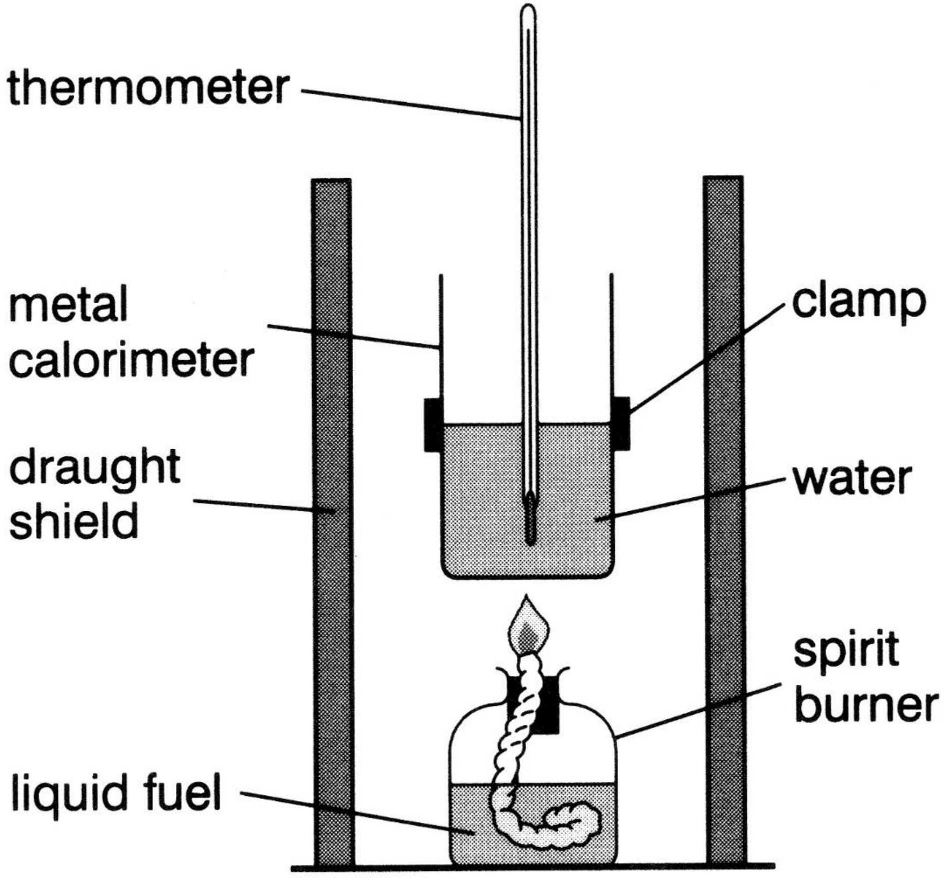
Experimental Design

Apparatus and Reagents

- Copper calorimeter
- Digital weighing machine
- Draught shield
- Glass rod
- Matches
- 100 cm^3 Measuring cylinder
- Retort stand and clamp
- Spirit burner and lid
- Thermometer

- Methanol
- Ethanol
- Propan-1-ol
- Water

Diagram



Experimental Design

Method

1) Weigh the copper calorimeter on the digital weighing machine. Record the data in the results table.

2) Use the measuring cylinder to pour 80 cm^3 of water into the copper calorimeter.

3) Re-weigh the copper calorimeter and water and hence calculate the mass of water present in the calorimeter. Record the data in the results table.

4) Support the copper calorimeter in the retort stand and clamp.

5) Place a thermometer in the copper calorimeter and measure the initial temperature of the water. Record the data in the results table.

6) Weigh the spirit burner and lid containing the methanol on the digital weighing machine. Record the data in the results table.

Experimental Design

Method

7) Place the spirit burner under the copper calorimeter. Adjust the height of the calorimeter so that it is 2 – 3 cm above the spirit burner.

8) Remove the lid from the spirit burner and, using the matches, immediately light the spirit burner. Adjust the height of the copper calorimeter once more so that the flame of the burning methanol covers the base of the calorimeter.

9) Using the glass rod, gently stir the water in the copper calorimeter. Ensure that no water spills out of the calorimeter.

10) When the temperature of the water in the copper calorimeter has increased by 20 °C, place the lid back on the spirit burner to extinguish the flame.

Experimental Design

Method

11) Measure the temperature of the water in the copper calorimeter at the end of the experiment. Record the data in the results table.

12) Calculate the change in temperature of the water in the calorimeter. Record the data in the results table.

13) Re-weigh the spirit burner on the digital weighing machine at the end of the experiment. Record the data in the results table.

14) Calculate the mass of methanol burned during the experiment. Record the data in the results table.

15) Repeat the experiment for the other two alcohols, ethanol and propan-1-ol.

16) Repeat the entire experiment at least once more.

Experimental Design

Results

	Methanol	Ethanol	Propan-1-ol
Mass of the copper calorimeter / g	X	X	X
Mass of the copper calorimeter + water / g	X	X	X
Mass of water in the copper calorimeter / g	X	X	X
Mass of spirit burner + lid + alcohol at the start of the reaction / g	X	X	X
Mass of spirit burner + lid + alcohol at the end of the reaction / g	X	X	X
Mass of alcohol burned during the reaction / g	X	X	X
Temperature of the water in the copper calorimeter at the start of the reaction / g	X	X	X
Temperature of the water in the copper calorimeter at the end of the reaction / g	X	X	X
Temperature increase of the water in the copper calorimeter during the reaction / g	X	X	X

Experimental Design

Calculation

- It is assumed that all of the energy released during combustion of the alcohol is absorbed by the water in the copper calorimeter, *i.e.*

energy produced by combustion of alcohol = energy gained by water

- Energy gained by water = $m_w \times c \times \Delta T$

where:

m_w = mass of water / g

c = specific heat capacity of water = 4.20 j / g / °C

ΔT = change in temperature of the water / °C

- Energy released by burning 1 g of alcohol = $(m_w \times c \times \Delta T) \div m_a$

where:

m_a = mass of alcohol / g

Experimental Design

Source of Error

1) The alcohol undergoes incomplete combustion.

Incomplete combustion of the alcohol releases less energy than complete combustion. This will reduce the amount by which the water in the calorimeter increases in temperature, resulting in a smaller value for the calculated energy change.

2) Heat energy is lost to the surroundings.

Heat energy from combustion of the alcohol is lost to the surroundings rather than being transferred into the water. This will reduce the amount by which the water in the calorimeter increases in temperature, resulting in a smaller value for the calculated energy change.

Experimental Design

Source of Error

3) Heat energy is absorbed by the copper calorimeter.

Heat energy from combustion of the alcohol is absorbed by the copper calorimeter rather than being transferred into the water. This will reduce the amount by which the water in the calorimeter increases in temperature, resulting in a smaller value for the calculated energy change.

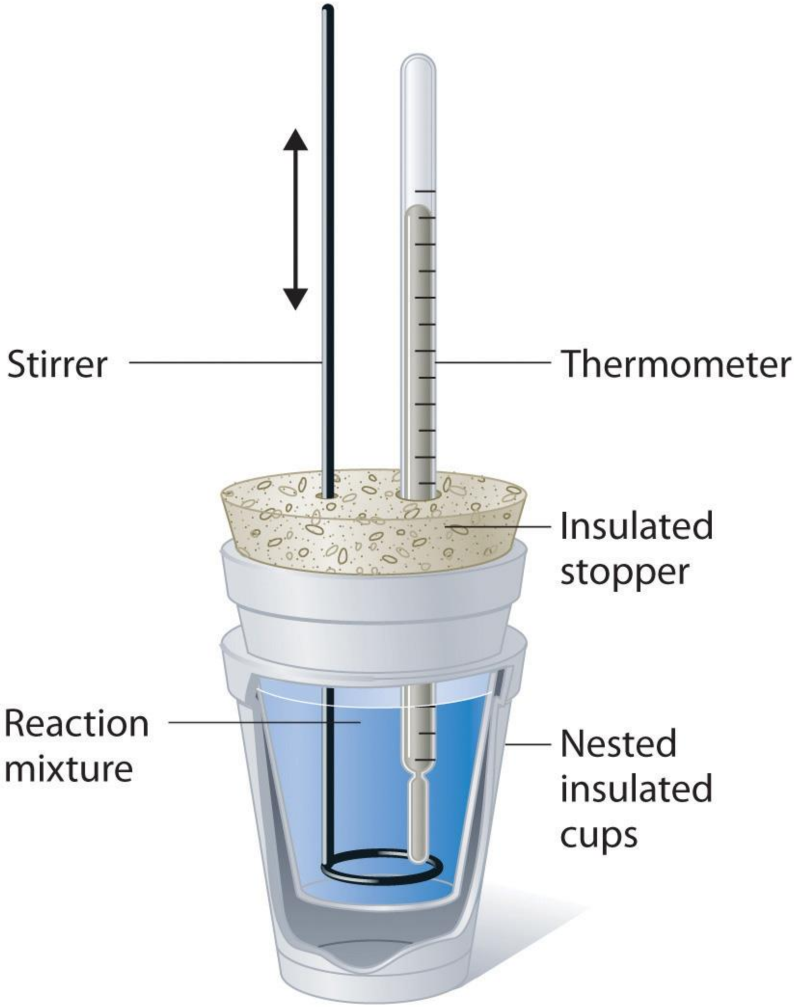
Experimental Design

- When ammonium chloride dissolves in water, an endothermic energy change takes place:

$NH_4Cl_{(s)} \rightarrow NH_4Cl_{(aq)} \quad +\Delta H$

- Design an experiment to determine the energy change when one mole of ammonium chloride dissolves completely in water.

Experimental Design



- The apparatus that could be used for conducting this experiment is shown in the diagram on the left.

- Refer to the diagram to help you design the experiment.

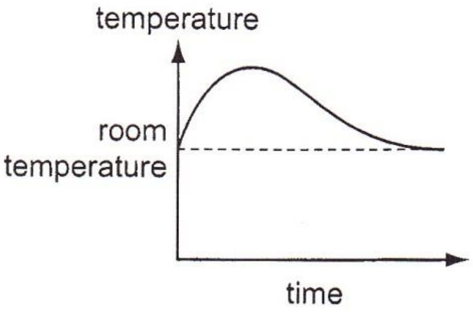
Energy from Chemicals

Could I please have some questions to *test my understanding* of the topic Energy from Chemicals?

Assessment for Learning
Question 1:

- Dissolving ammonium nitrate in water is endothermic. Which graph shows how the temperature alters as the ammonium nitrate is added to water and the solution is left to stand?

A



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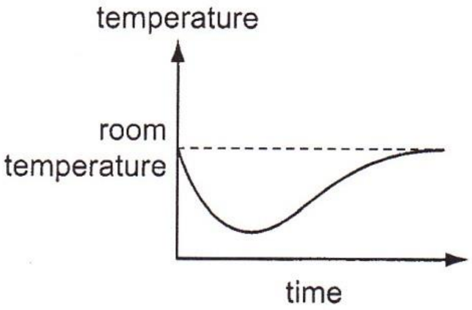

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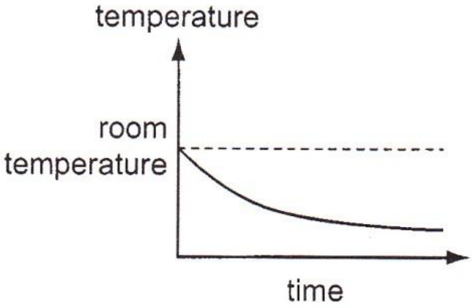

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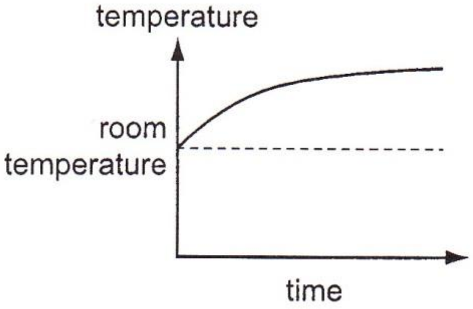

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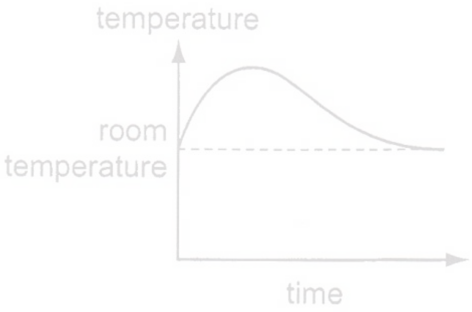

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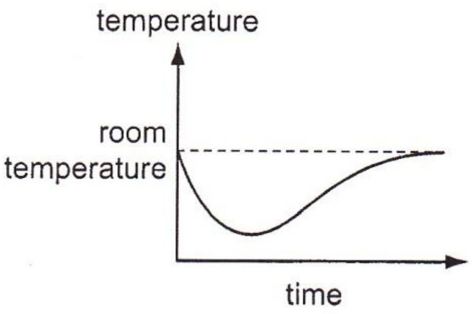

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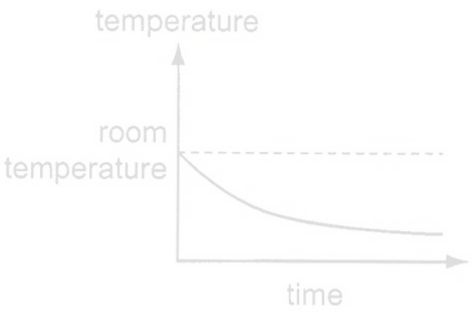

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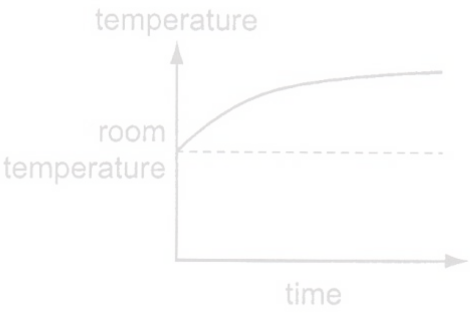

Assessment for Learning Question 2:

- The formation of hydrogen iodide from hydrogen and iodine is an endothermic reaction.

$$H{-}H + I{-}I \rightarrow H{-}I + H{-}I$$

What may be deduced from this information?

A The number of bonds broken is greater than the number of bonds formed.

B The formation of H–I bonds absorbs energy.

C The products possess less energy than the reactants.

D The total energy change in bond formation is less than that in bond breaking.

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Assessment for Learning Question 3:

- The table below shows the energy released by the complete combustion of some compounds used as fuels.

Compound	Formula	M_r	ΔH in kJ / mol
methane	CH_4	16	–880
ethanol	C_2H_5OH	46	–1380
propane	C_3H_8	44	–2200
heptane	C_7H_{16}	100	–4800

Which fuel produces the most energy when 1 g of the compound is completely burned?

A Ethanol

B Heptane

C Methane

D Propane

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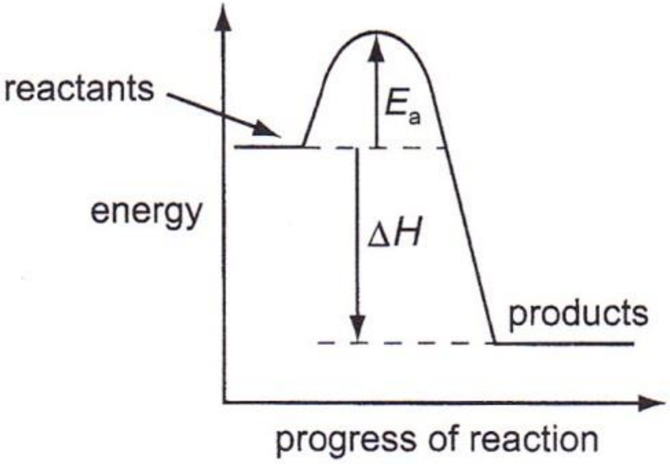
Assessment for Learning Question 4:

- The enthalpy diagram shows an uncatalysed, exothermic reaction.



The reaction was repeated in the presence of a catalyst. What effect does the catalyst have on the activation energy , E_a, and the enthalpy change, ΔH?

A E_a decreases and ΔH decreases.
B E_a decreases and ΔH unchanged.
C E_a increases and ΔH increases.
D E_a unchanged and ΔH decreases.

Assessment for Learning
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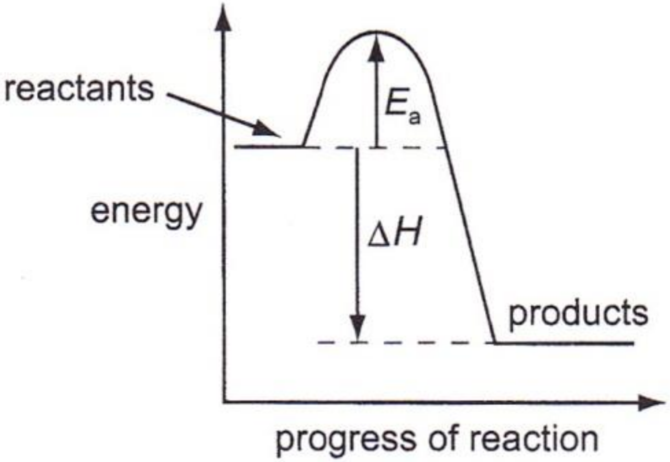

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Assessment for Learning
Question 5:

The scheme shows four stages, I to IV, in the conversion of solid candlewax, $C_{30}H_{62}$, into carbon dioxide and water.

I $C_{30}H_{62(s)} \rightarrow C_{30}H_{62(l)}$
II $C_{30}H_{62(l)} \rightarrow C_{30}H_{62(g)}$
III $C_{30}H_{62(g)} + 45.5O_{2(g)} \rightarrow 30CO_{2(g)} + 31H_2O_{(g)}$
IV $30CO_{2(g)} + 31H_2O_{(g)} \rightarrow 30CO_{2(g)} + 31H_2O_{(l)}$

Which stages are exothermic?

A I and II.
B II and III.
C III and IV.
D I and IV.

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Presentation on
Energy from Chemicals
by Dr. Chris Slatter

christopher_john_slatter@nygh.edu.sg

Nanyang Girls' High School
2 Linden Drive
Singapore
288683

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