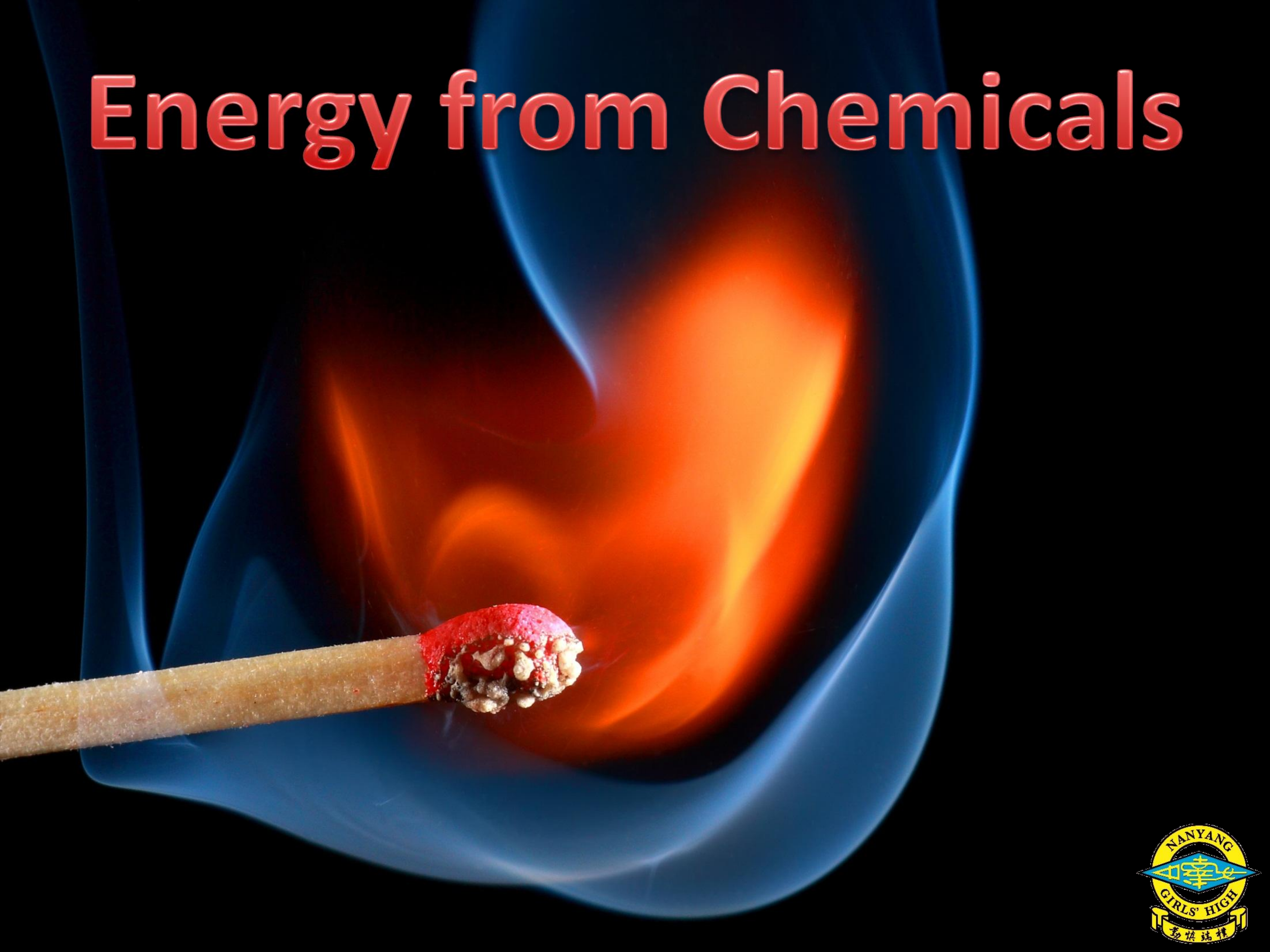
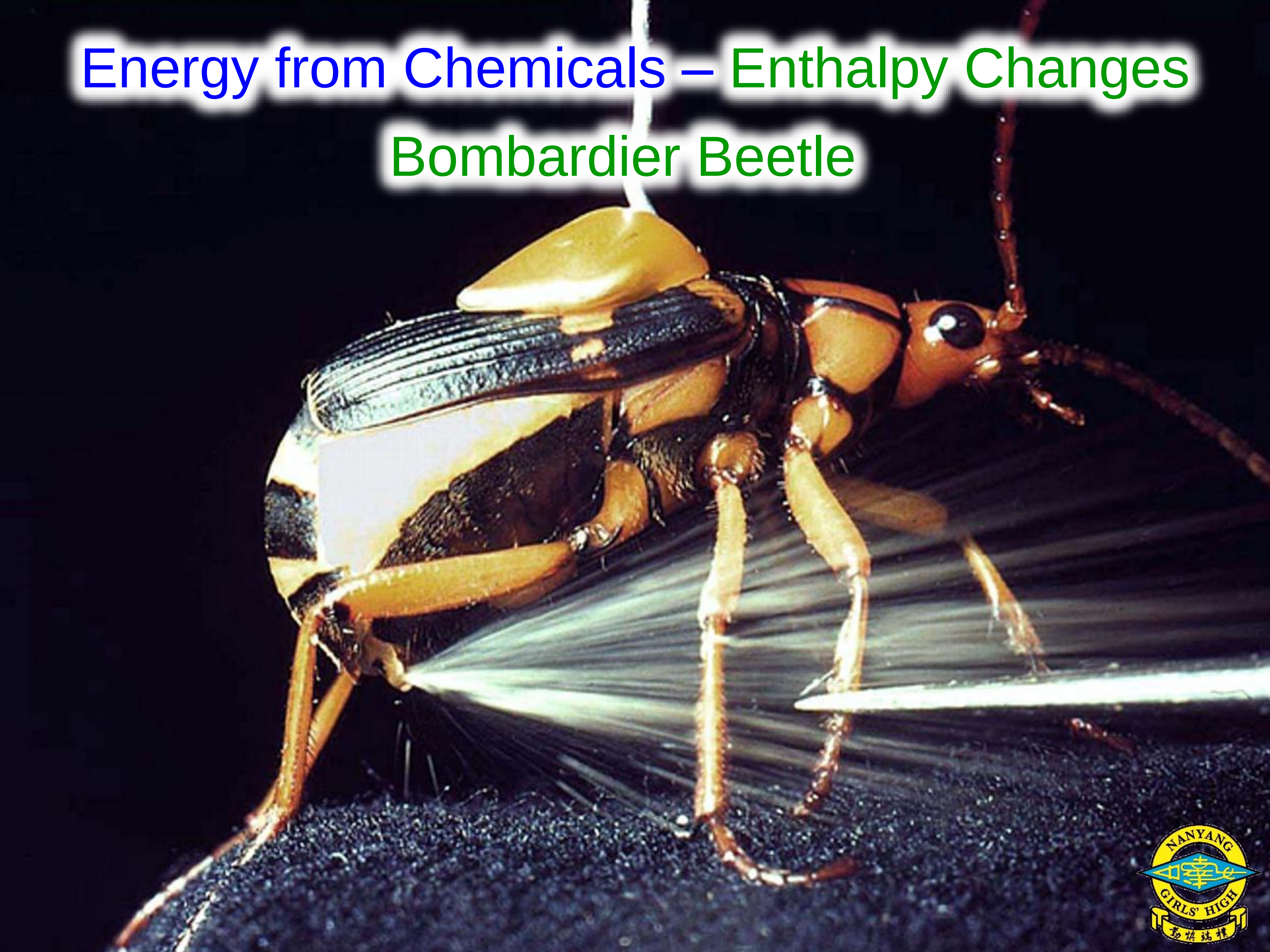
Energy from Chemicals



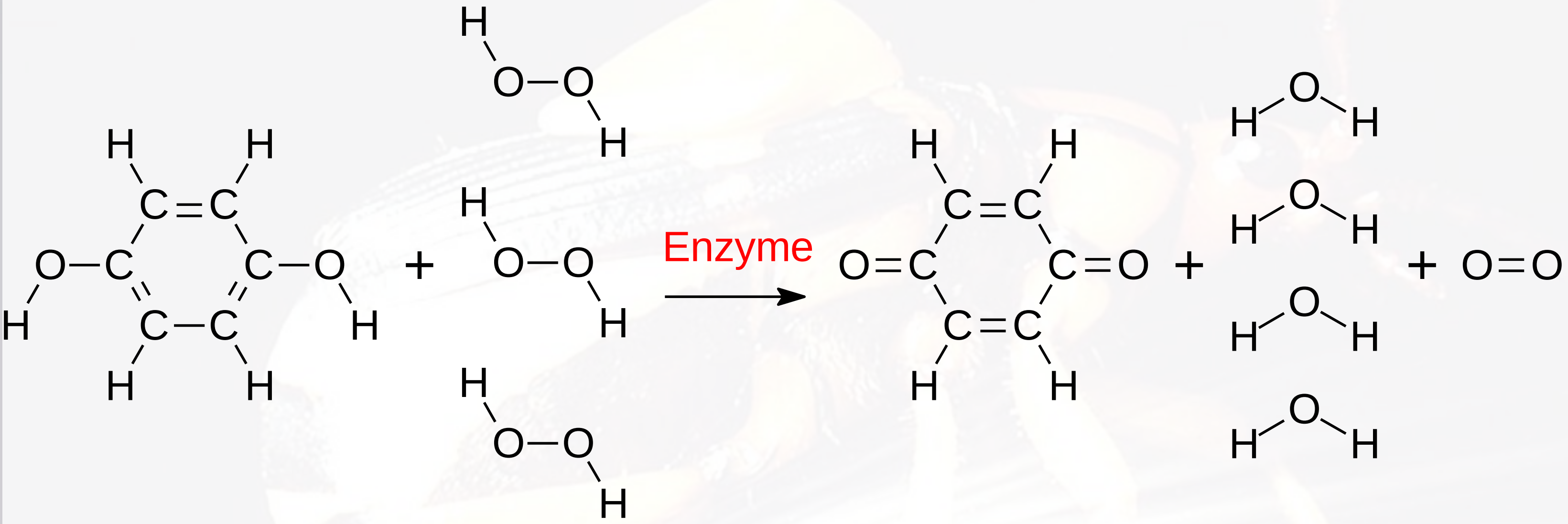
Energy from Chemicals – Enthalpy Changes

Bombardier Beetle



Energy from Chemicals – Enthalpy Changes

Bombardier Beetle

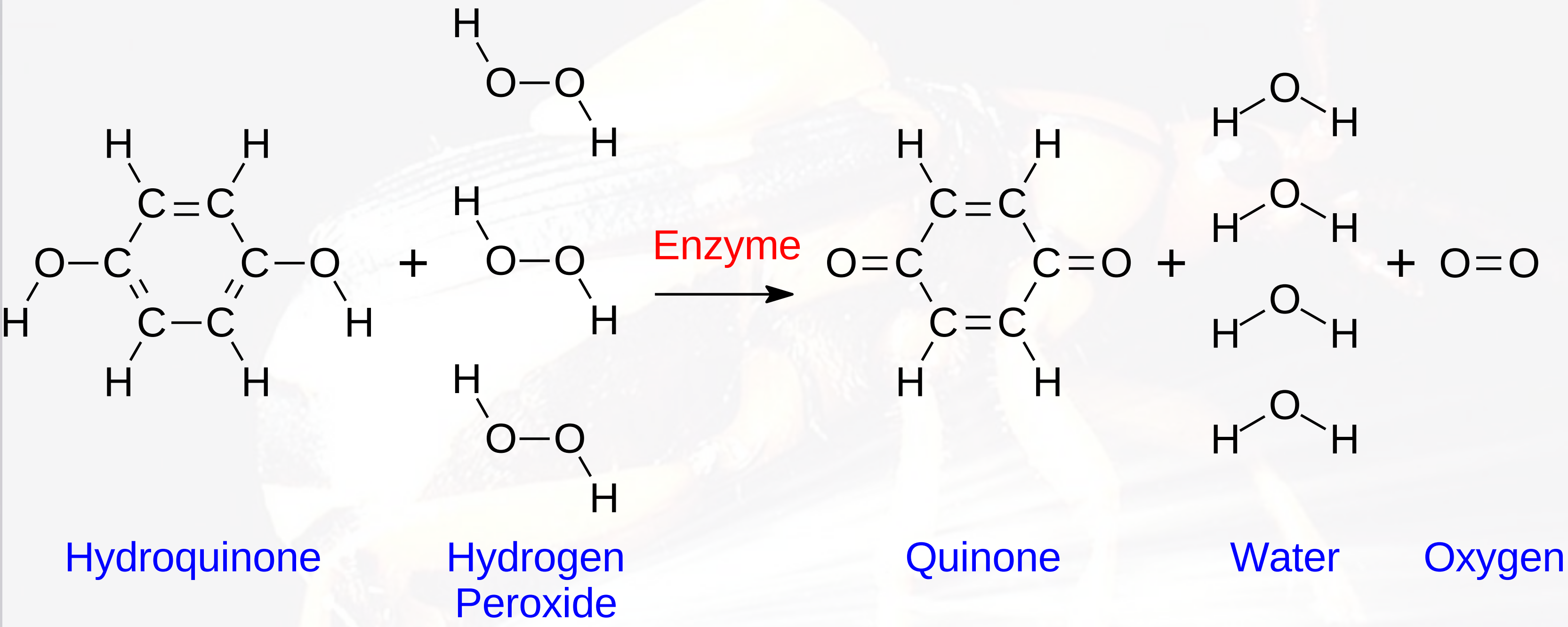


Hydroquinone + Hydrogen Peroxide → Quinone + Water + Oxygen

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

Energy from Chemicals – Enthalpy Changes

Bombardier Beetle



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

Energy from Chemicals – Enthalpy Changes

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

- What are some generalisations about *change*?

Energy from Chemicals – Enthalpy Changes

- 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 – Enthalpy Changes

- **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 – Enthalpy Changes

- **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 – Enthalpy Changes

What must I know and understand about *energy from chemicals*?

Energy from Chemicals – Enthalpy Changes

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 – Enthalpy Changes

Why do Chemists use the term *enthalpy* instead of *energy*?

Energy from Chemicals – Enthalpy Changes

- Chemists use thermodynamics to keep track of individual changes in energy, and to assess the net change that takes place in a reaction. For this purpose, they use an assessment of the quantity of energy available from a reaction as heat that is called the *enthalpy*. The name comes from the Greek words for “*heat inside*”. There are good technical grounds for distinguishing *enthalpy* from *energy*, but essentially, *enthalpy* can be considered as just another name for *energy trapped in chemicals and available as heat*.

Energy from Chemicals – Enthalpy Changes

What different *enthalpy changes* can take place in Chemistry?

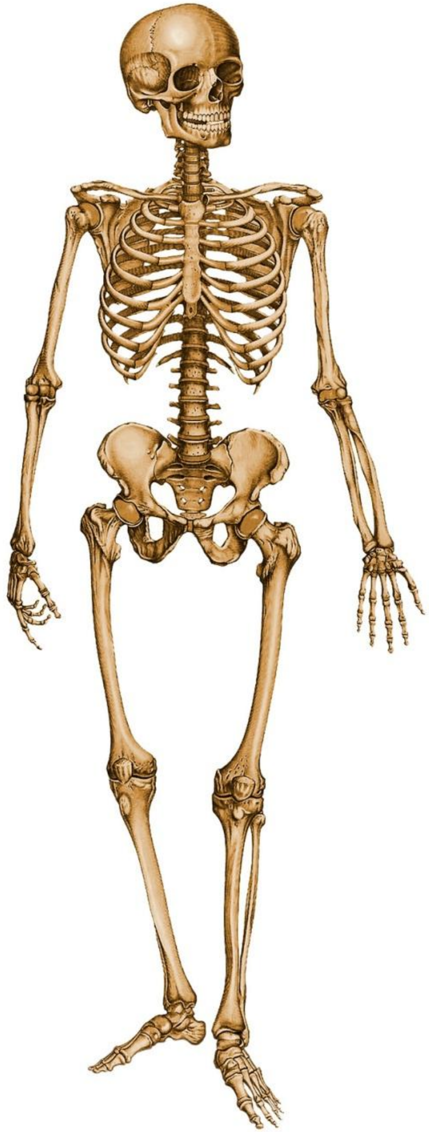
- Exothermic: Releases energy into the surroundings.
- Endothermic: Absorbs energy from the surroundings.

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



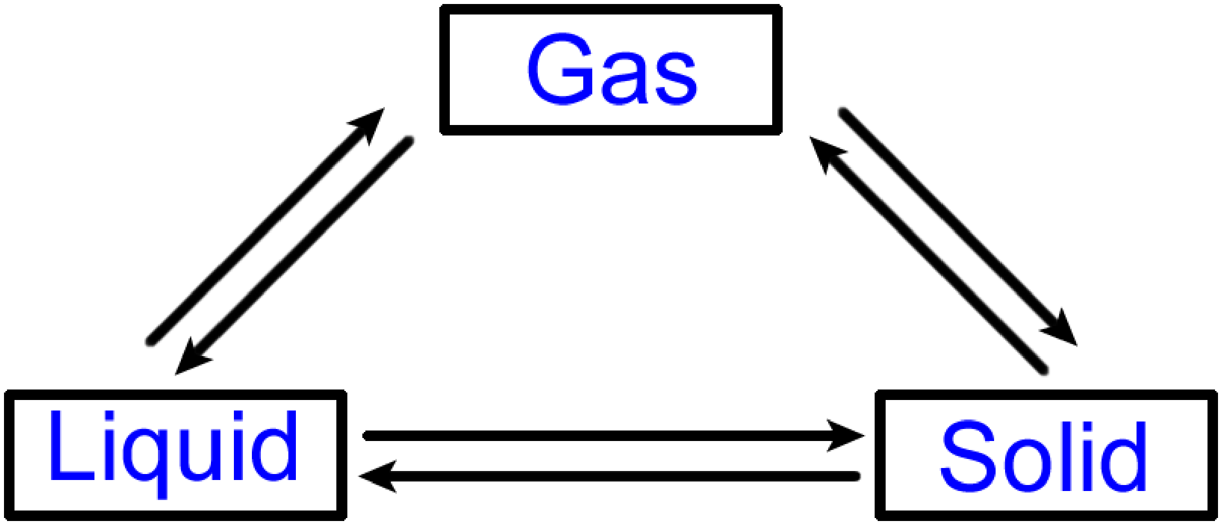
Endo is Greek for “inside.”
Thermasi is Greek for “to heat.”
The literal meaning of endothermic is therefore “to heat inside.”

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 – Enthalpy Changes

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

Energy from Chemicals – Enthalpy Changes

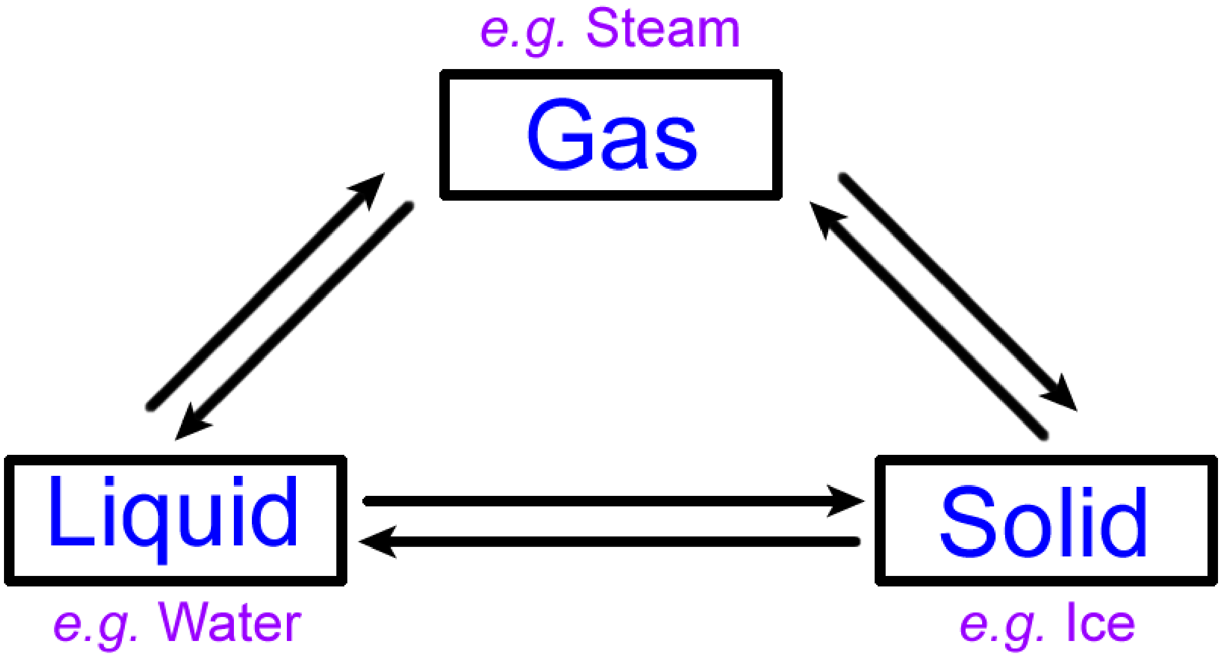


Gas

Liquid

Solid

Energy from Chemicals – Enthalpy Changes



e.g. Steam

Gas

Liquid

e.g. Water

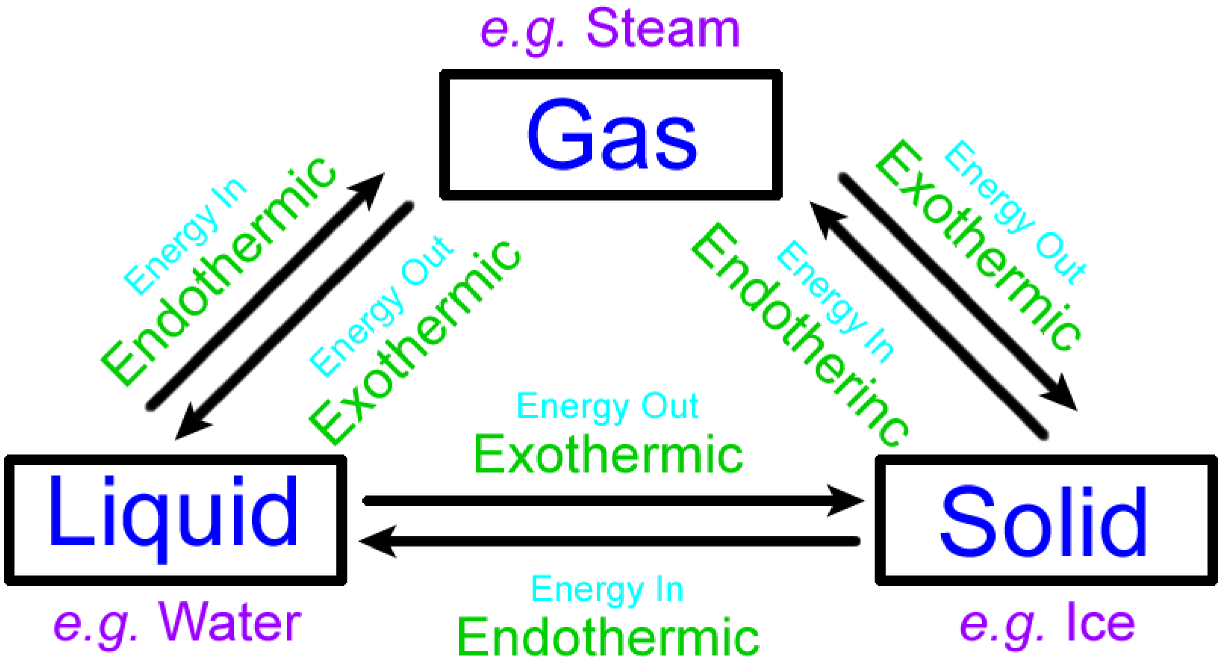
Solid

e.g. Ice

Remember:

energy $H_2O(g)$ > energy $H_2O(\mathit{l})$ > energy $H_2O(s)$

Energy from Chemicals – Enthalpy Changes



e.g. Steam

Gas

Energy In

Endothermic

Energy Out

Exothermic

Energy Out

Exothermic

Energy In

Endotherinc

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 – Enthalpy Changes

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 – Enthalpy Changes

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 Profile for an Exothermic Reaction

Please tell me
about *exothermic*
chemical reactions.

Exothermic Reactions

Energy Profile for an Exothermic Reaction

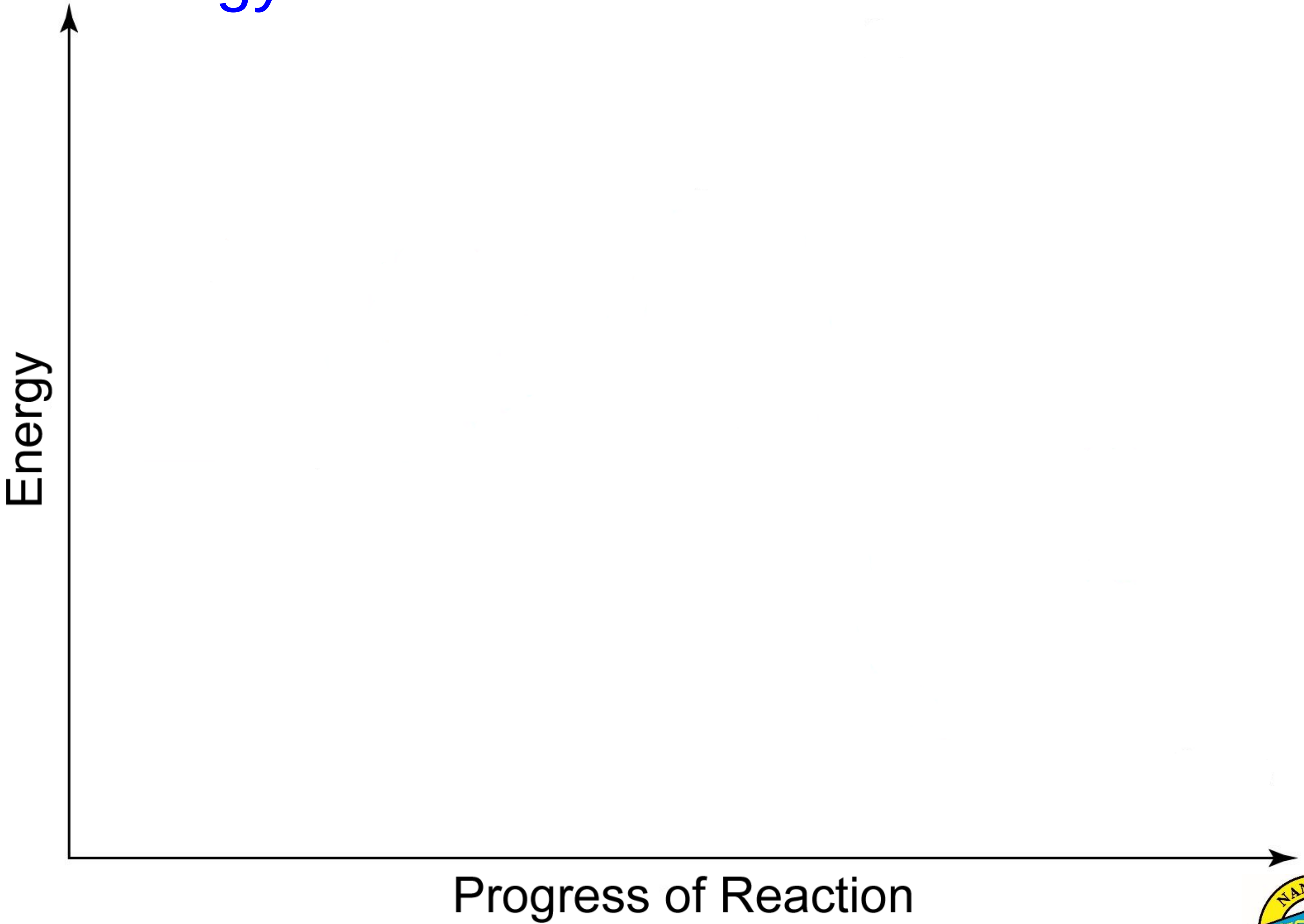
Important Things to Note…

- 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



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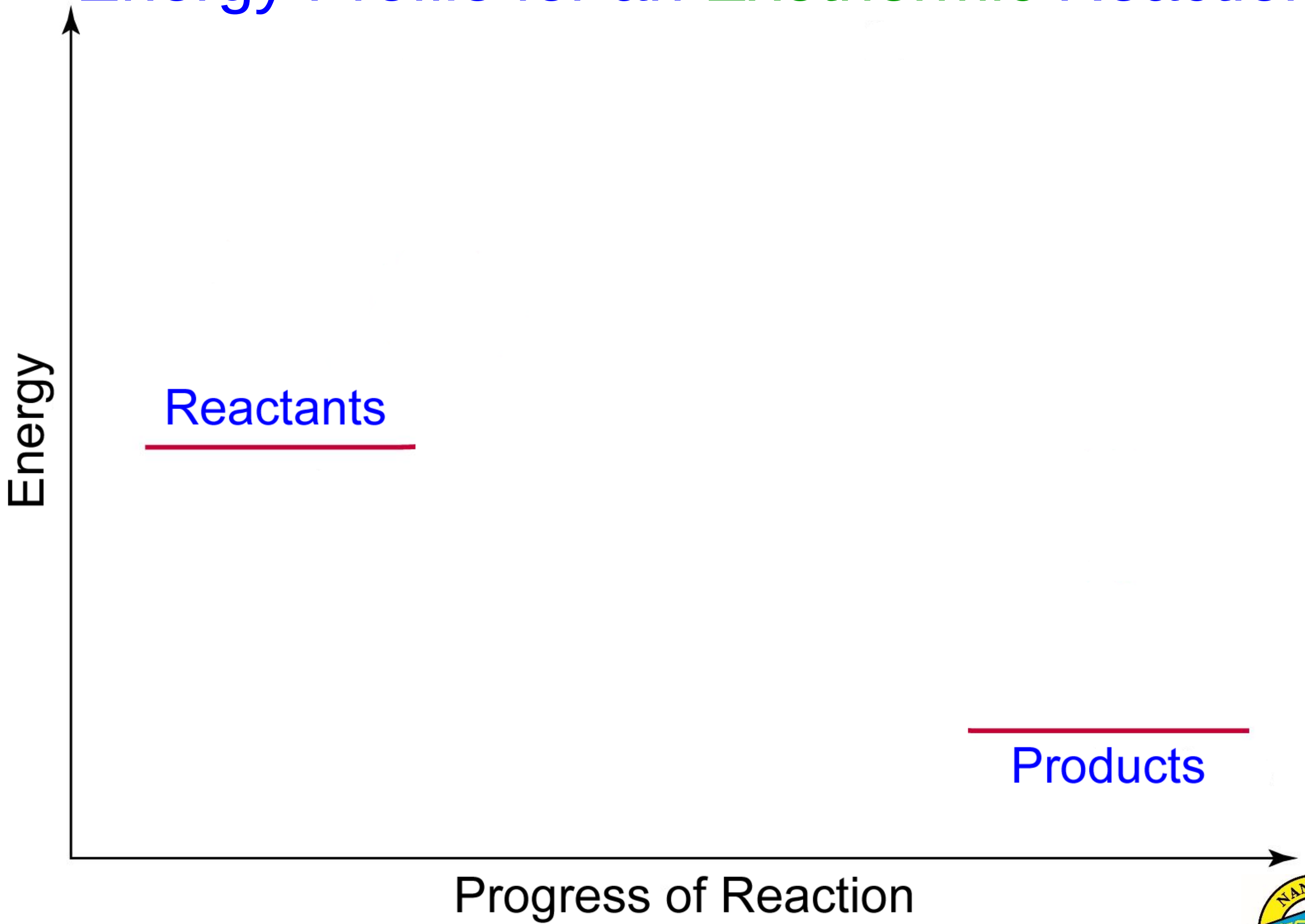

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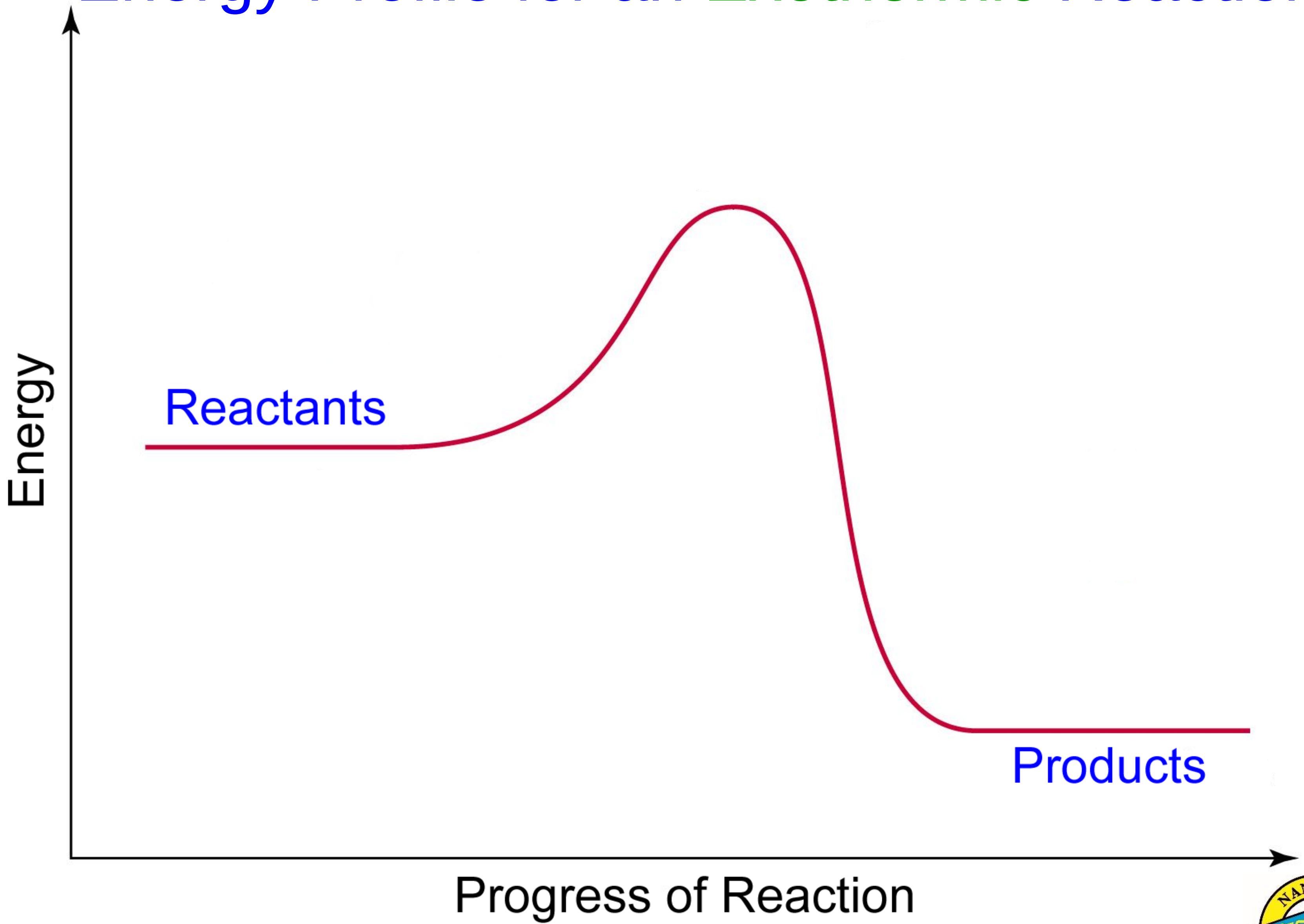
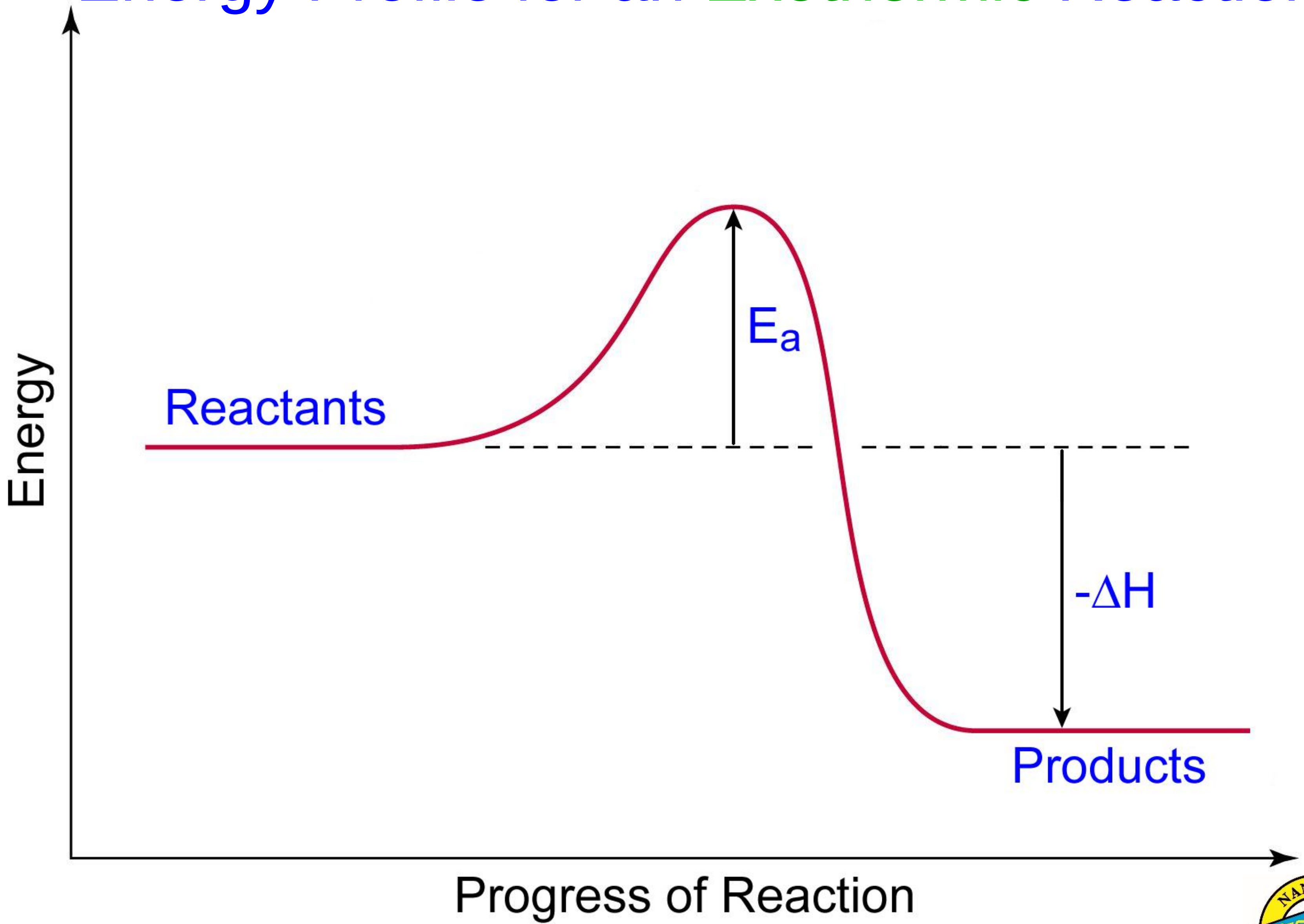


Energy Profile for an Exothermic Reaction



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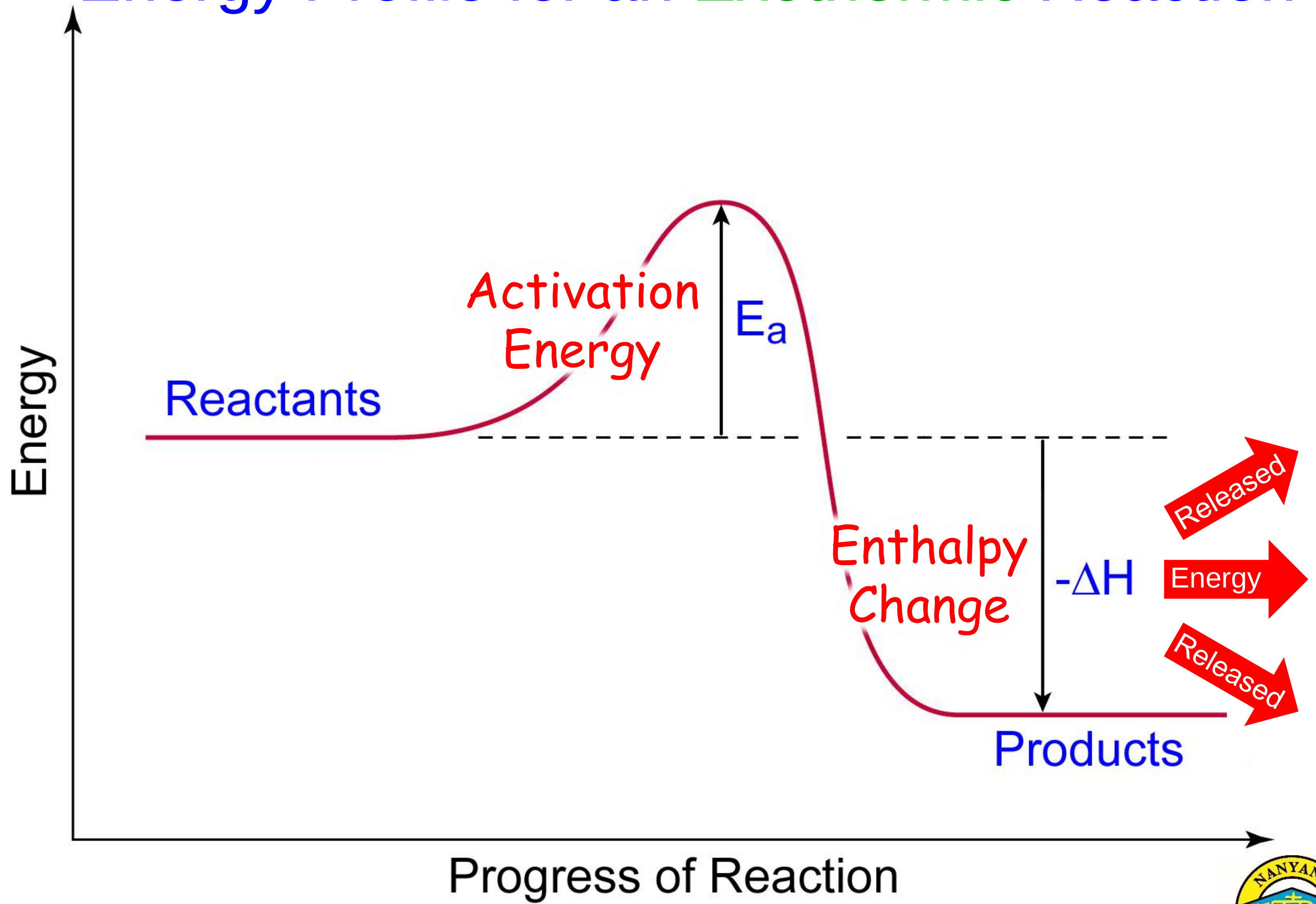
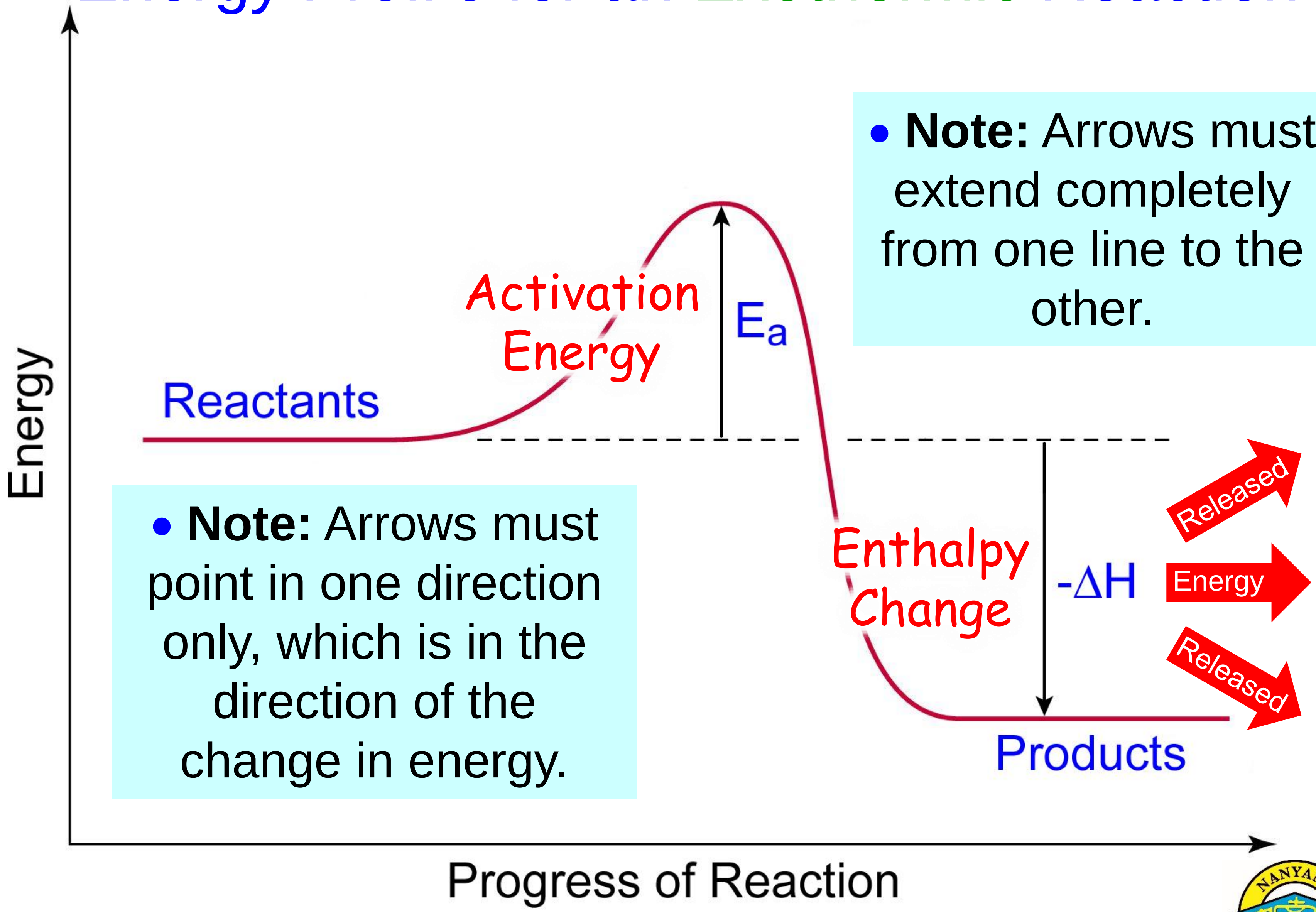


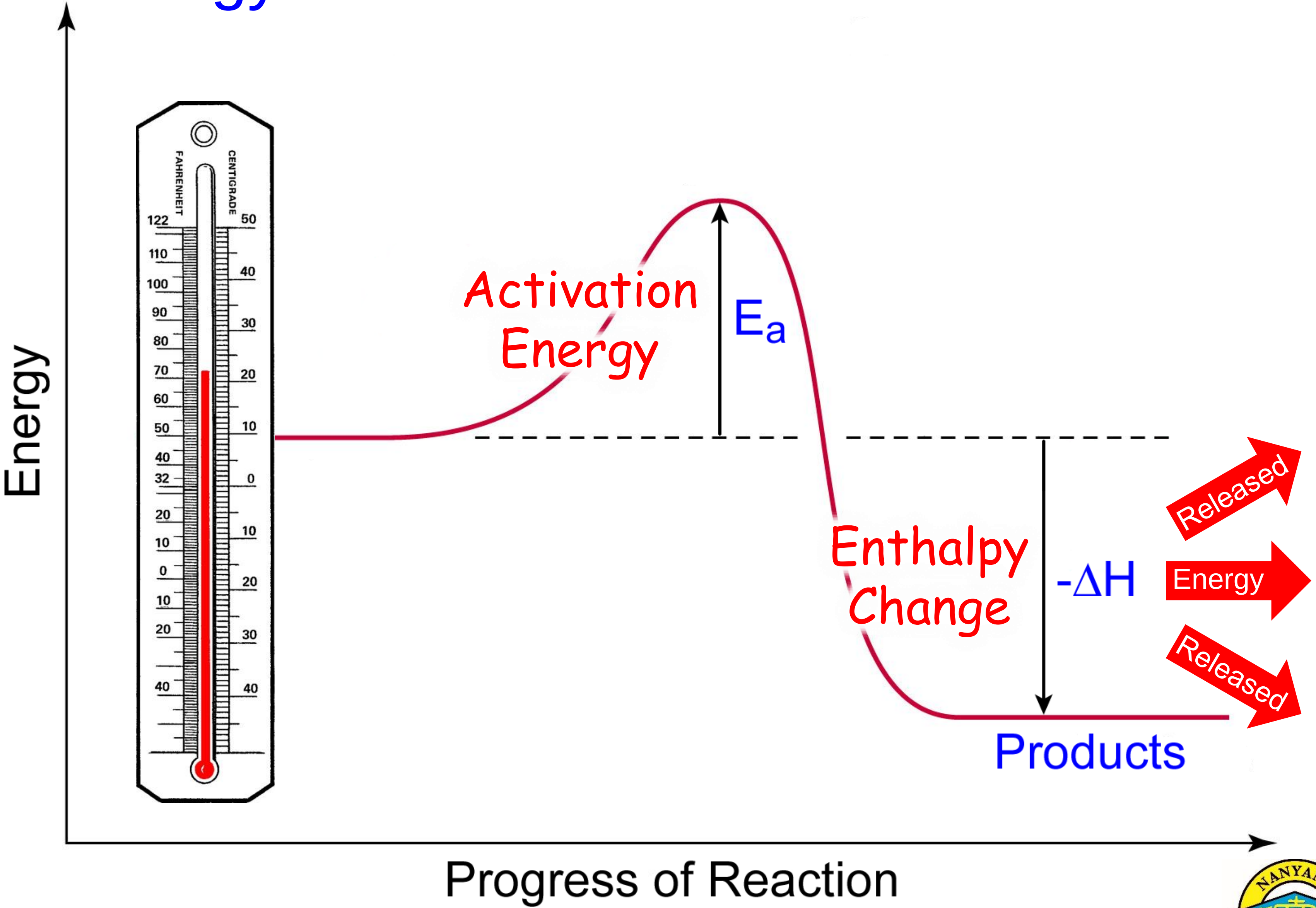
Energy Profile for an Exothermic Reaction



- **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



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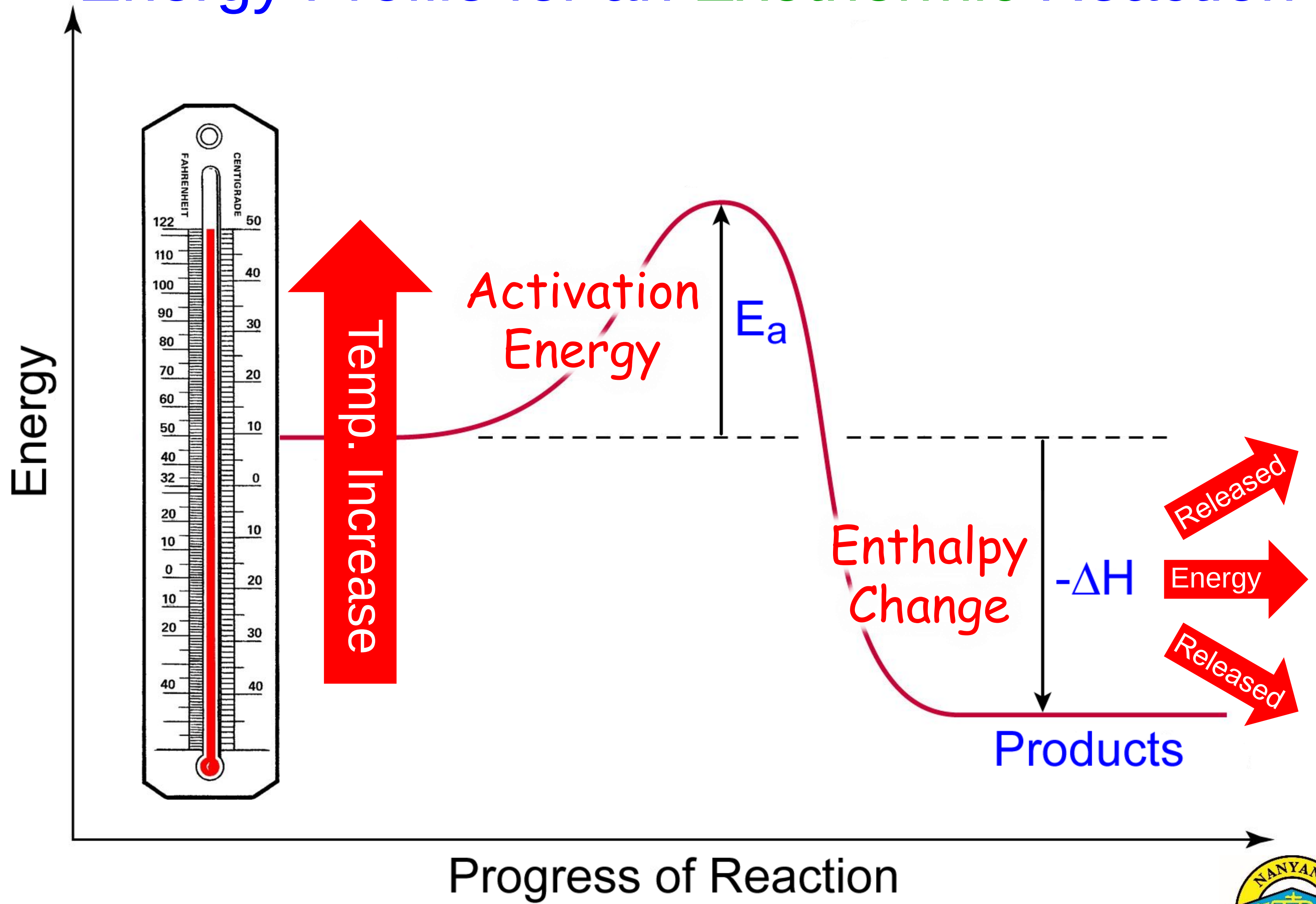

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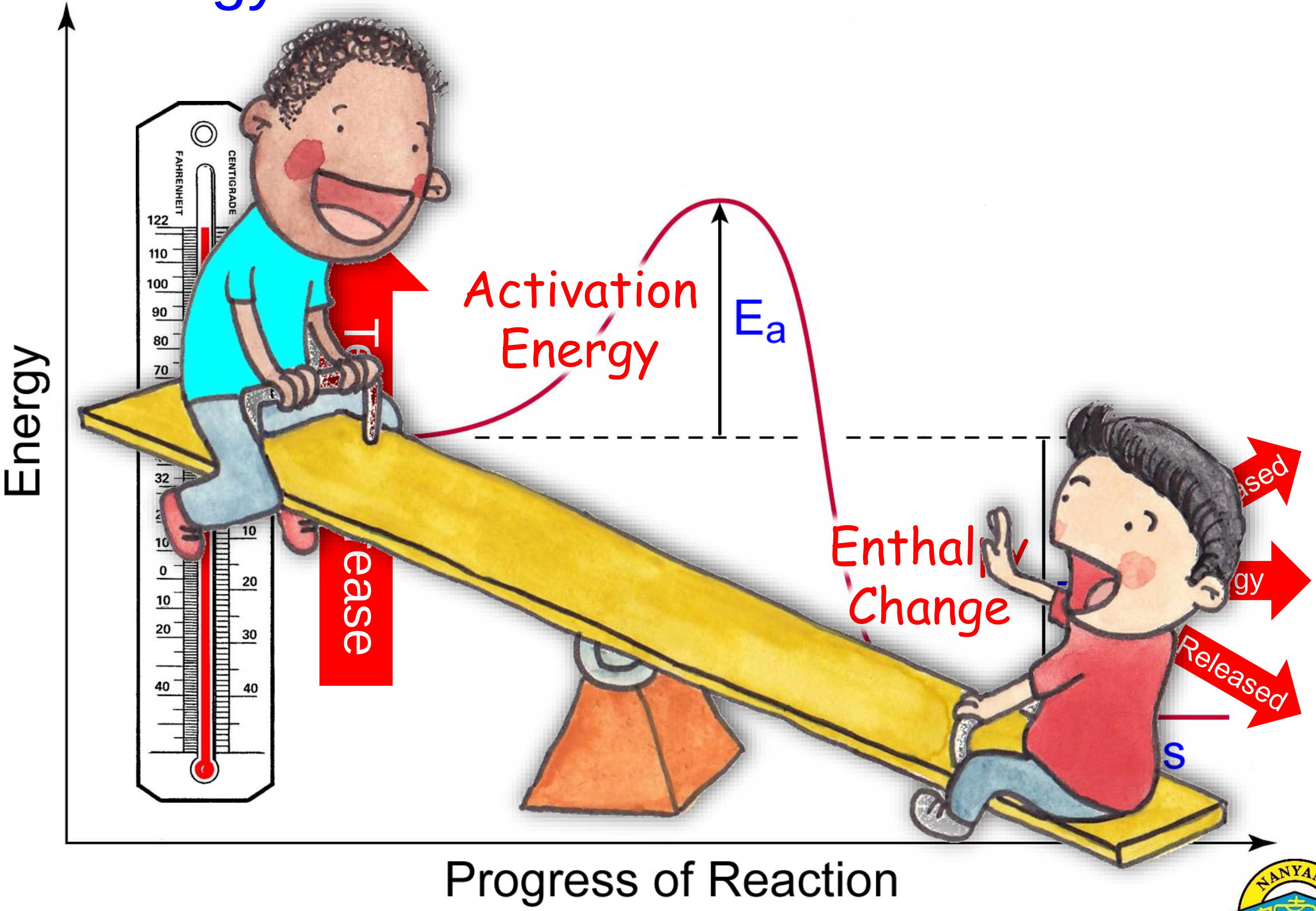


Energy Profile for an Exothermic Reaction

Energy Profile for an Exothermic Reaction

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

- For example, the combustion of methane:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

$$\Delta H = -890 \text{ kJ mol}^{-1}$$

- For example, the combustion of ethanol:

$$C_2H_5OH(g) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$$

$$\Delta H = -1371 \text{ kJ mol}^{-1}$$

Energy from Chemicals – Enthalpy Changes

- 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 – Enthalpy Changes

- 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.

NANYANG GIRLS' HIGH

Activation Energy

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.

Energy Profile for an Endothermic Reaction

Please tell me
about *endothermic*
chemical reactions.

Endothermic Reaction

Energy Profile for an Endothermic Reaction

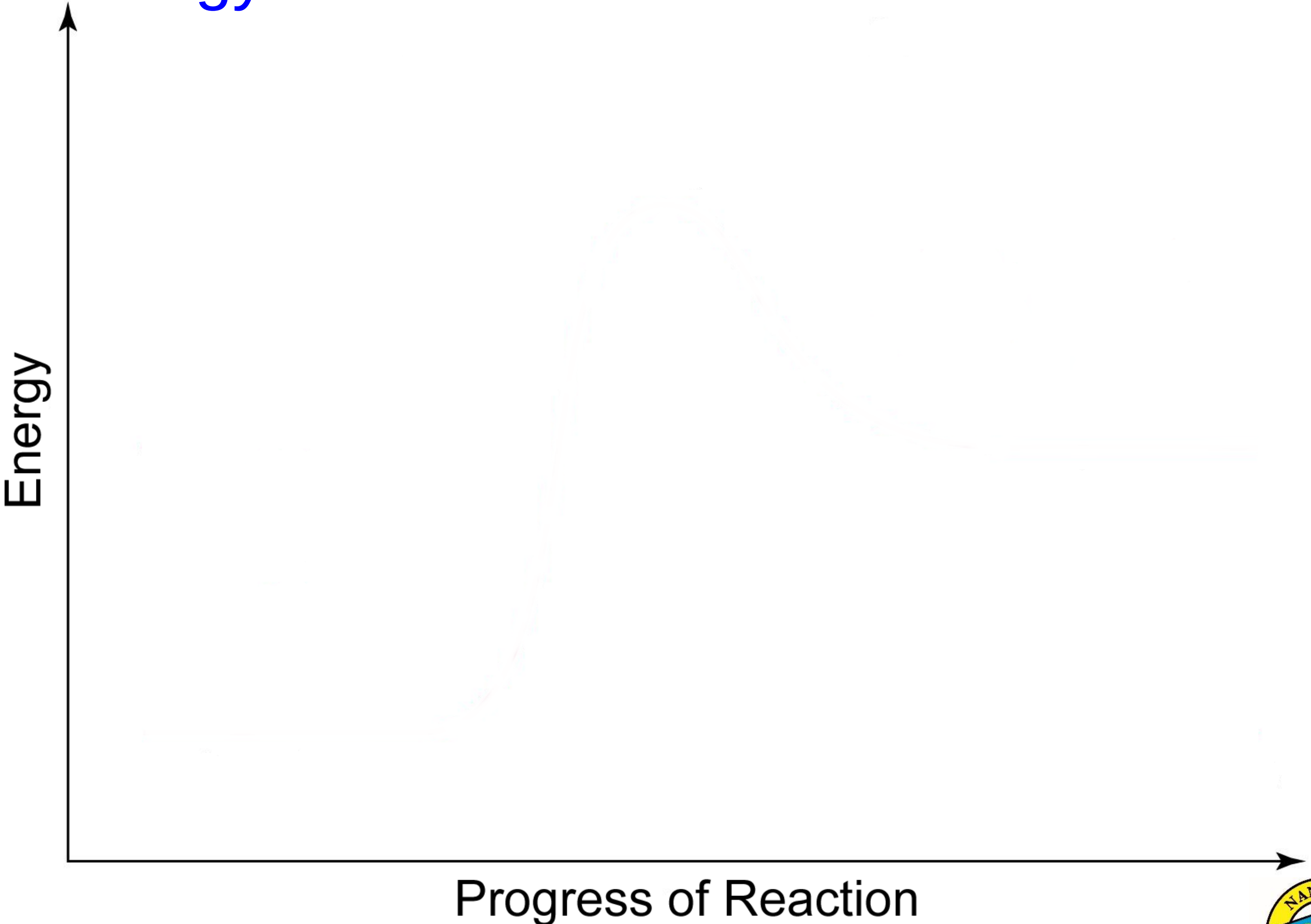
Important Things to Note…

- ΔH means change in enthalpy.

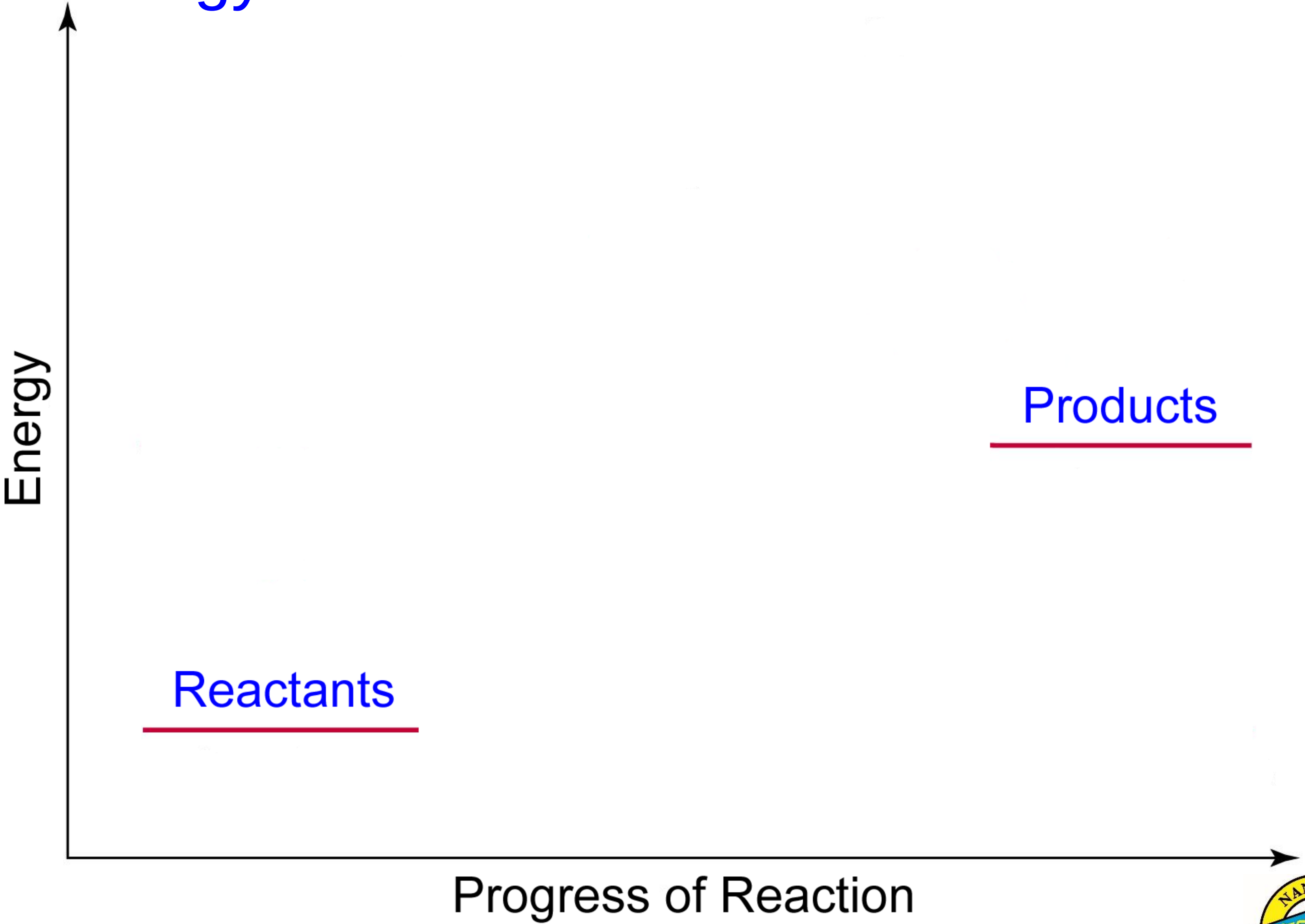
- An endothermic change absorbs energy from the surroundings. Numerical values of ΔH are positive.

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

Energy Profile for an Endothermic Reaction



Energy Profile for an Endothermic Reaction



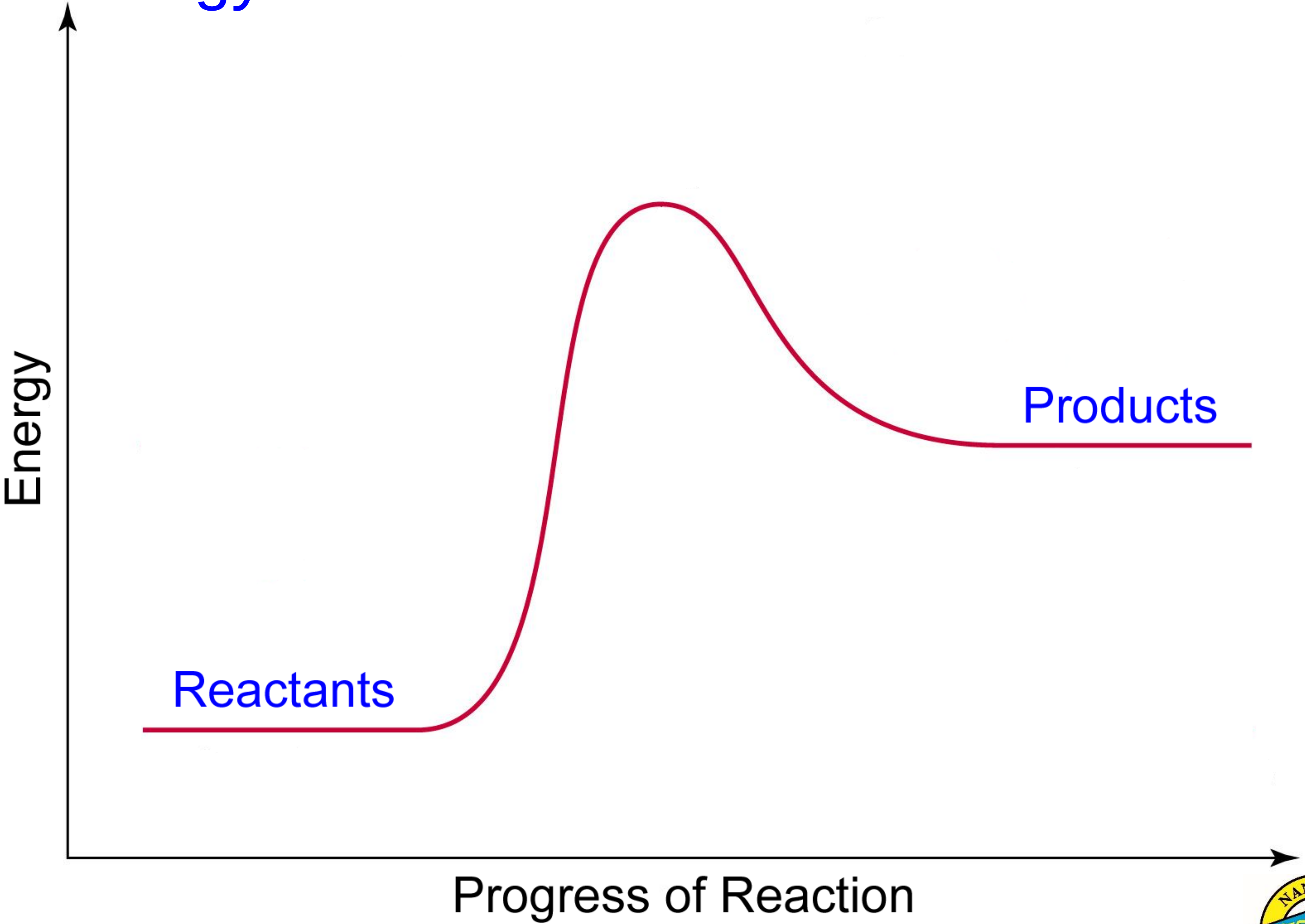
Energy

Products

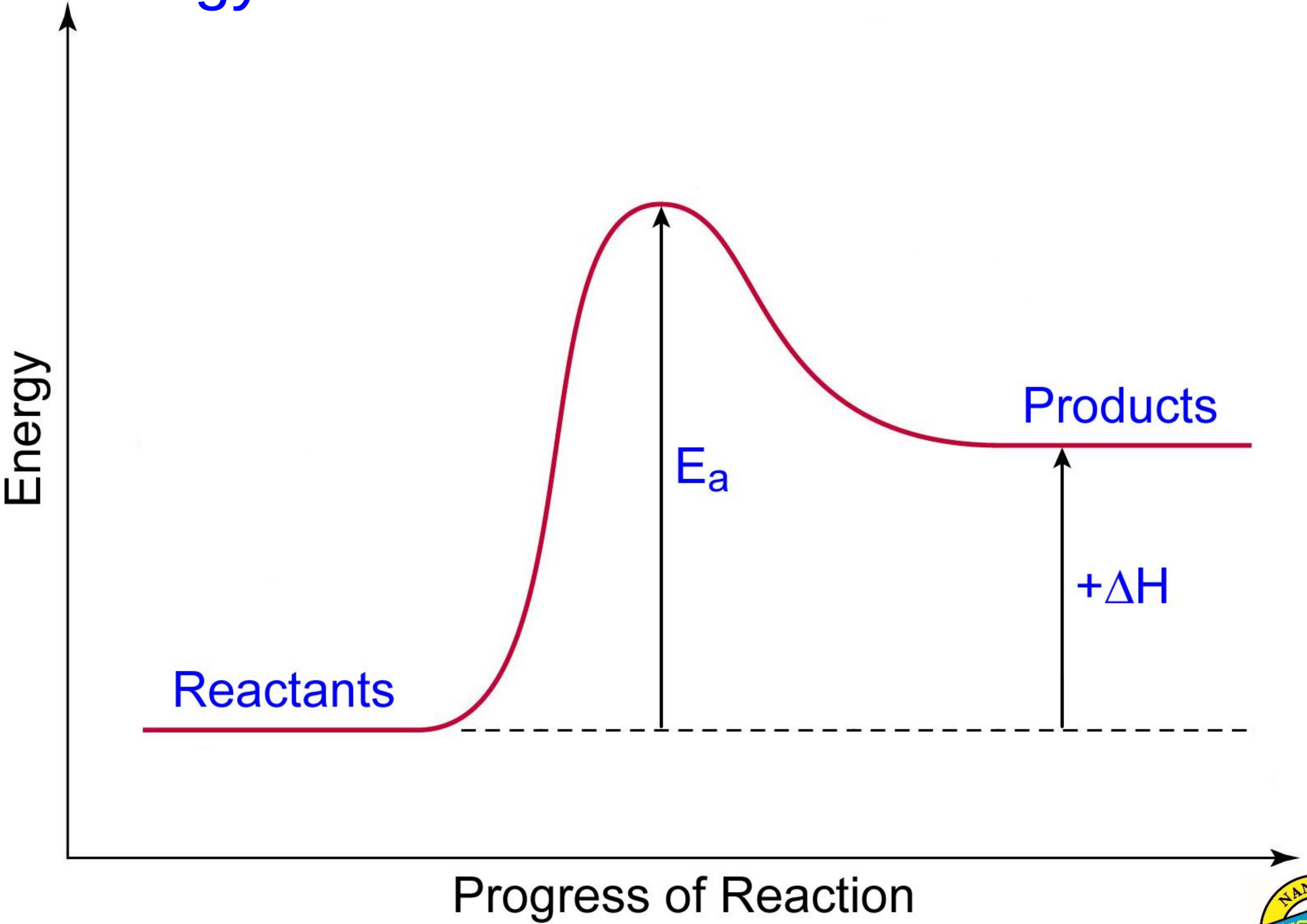
Reactants

Progress of Reaction

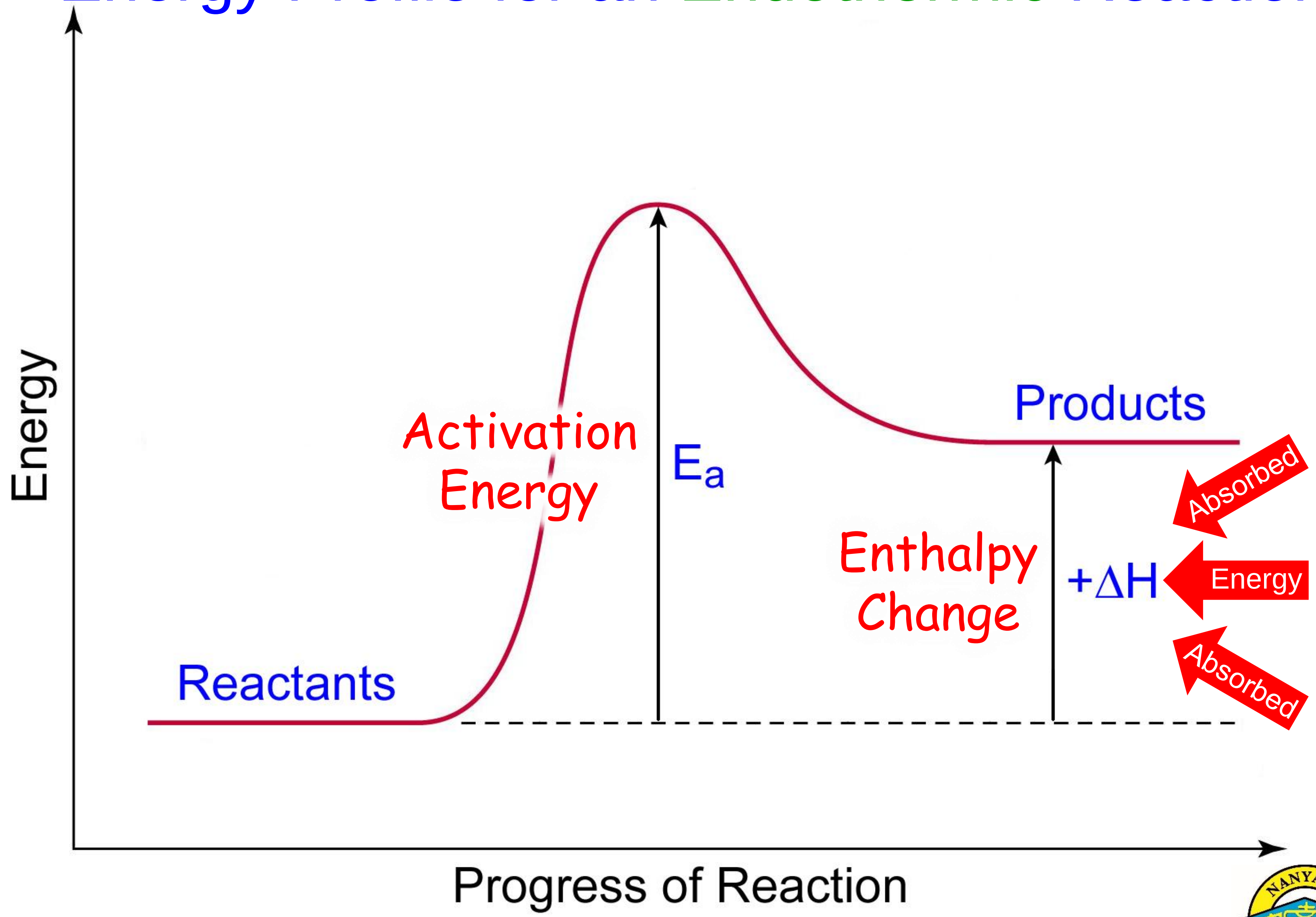
Energy Profile for an Endothermic Reaction



Energy Profile for an Endothermic Reaction



Energy Profile for an Endothermic Reaction



Energy Profile for an Endothermic Reaction

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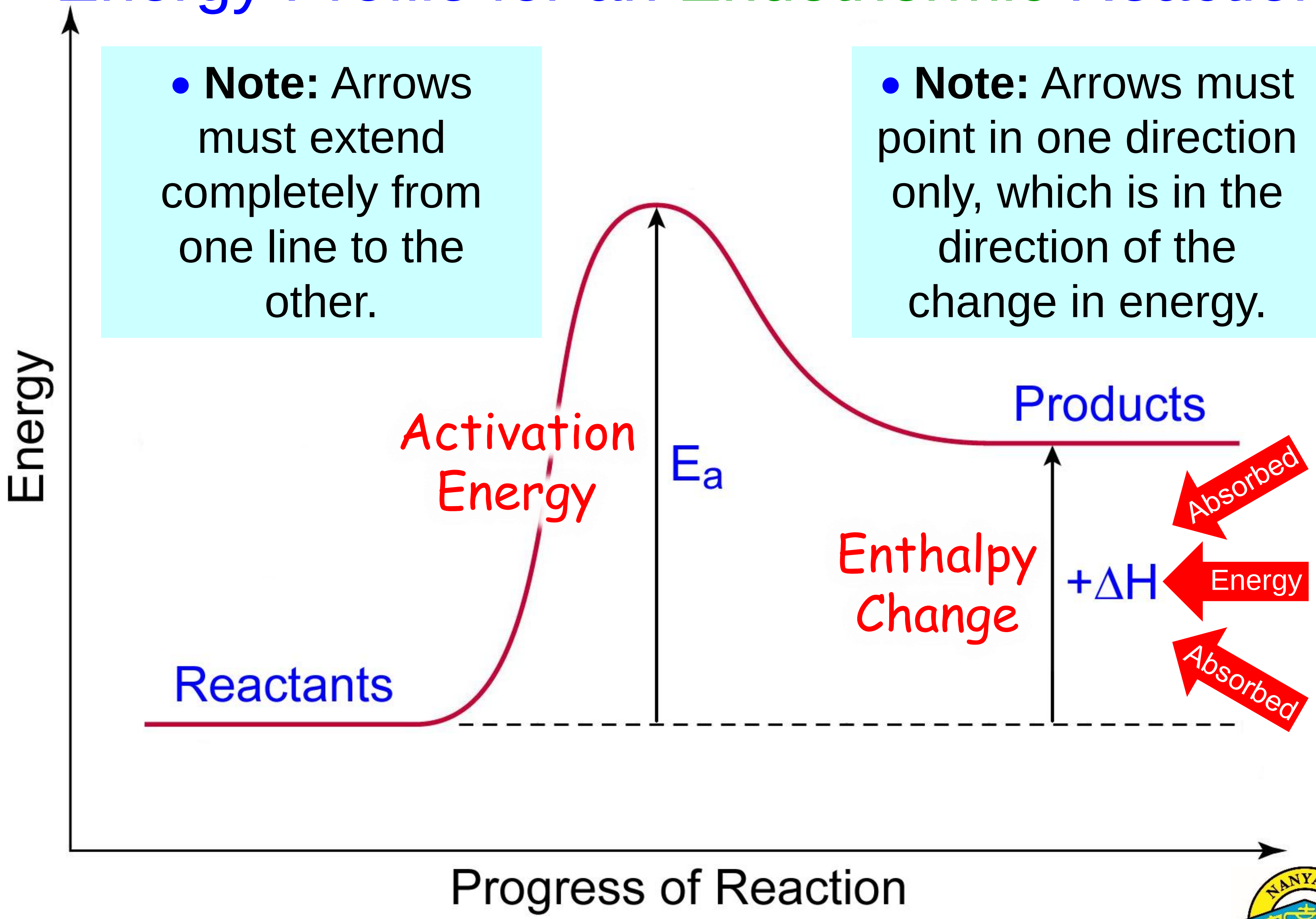
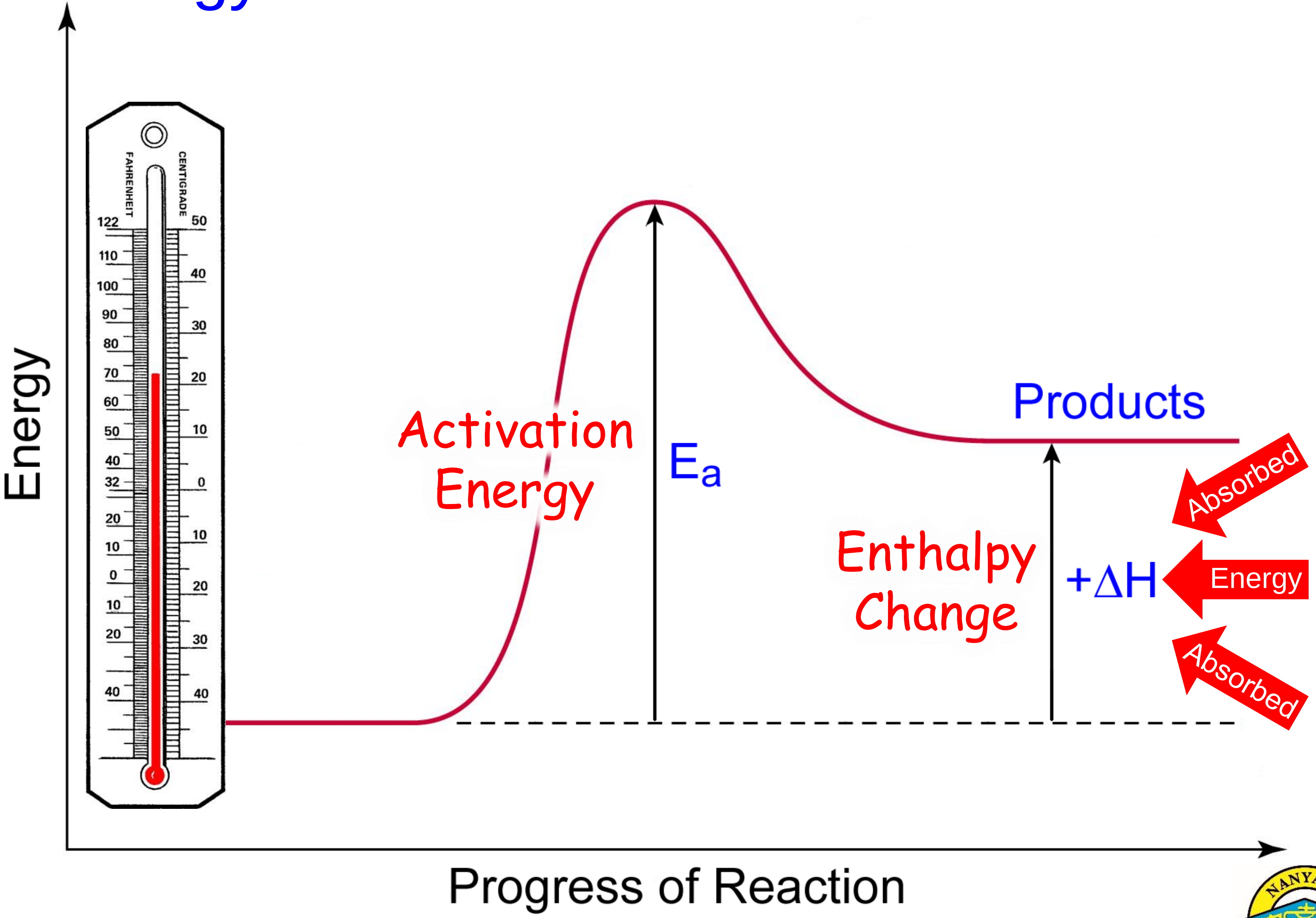
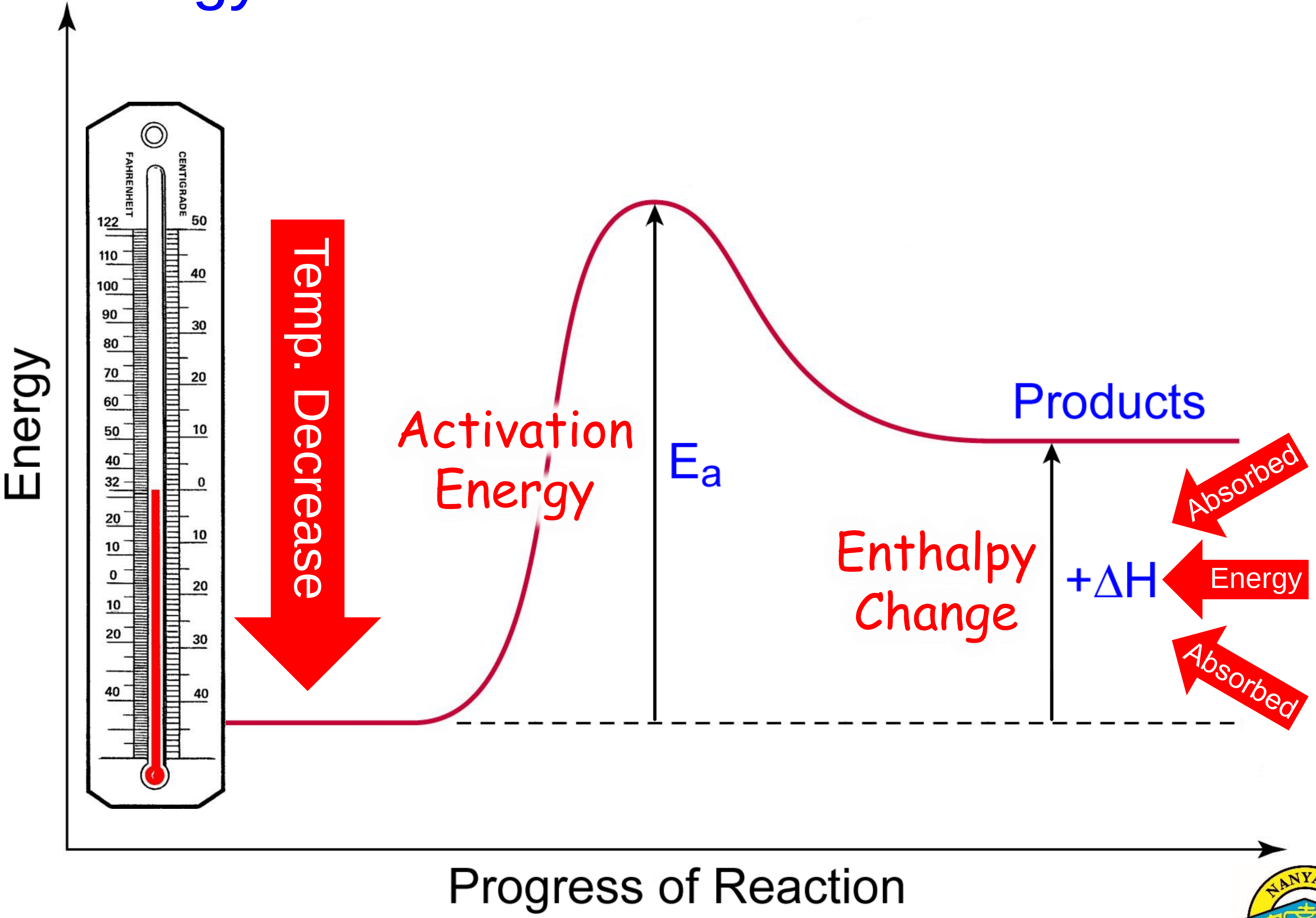


Energy Profile for an Endothermic Reaction



Energy Profile for an Endothermic Reaction



Temp. Decrease

Activation Energy

E_a

Products

Enthalpy Change

$+\Delta H$

Absorbed

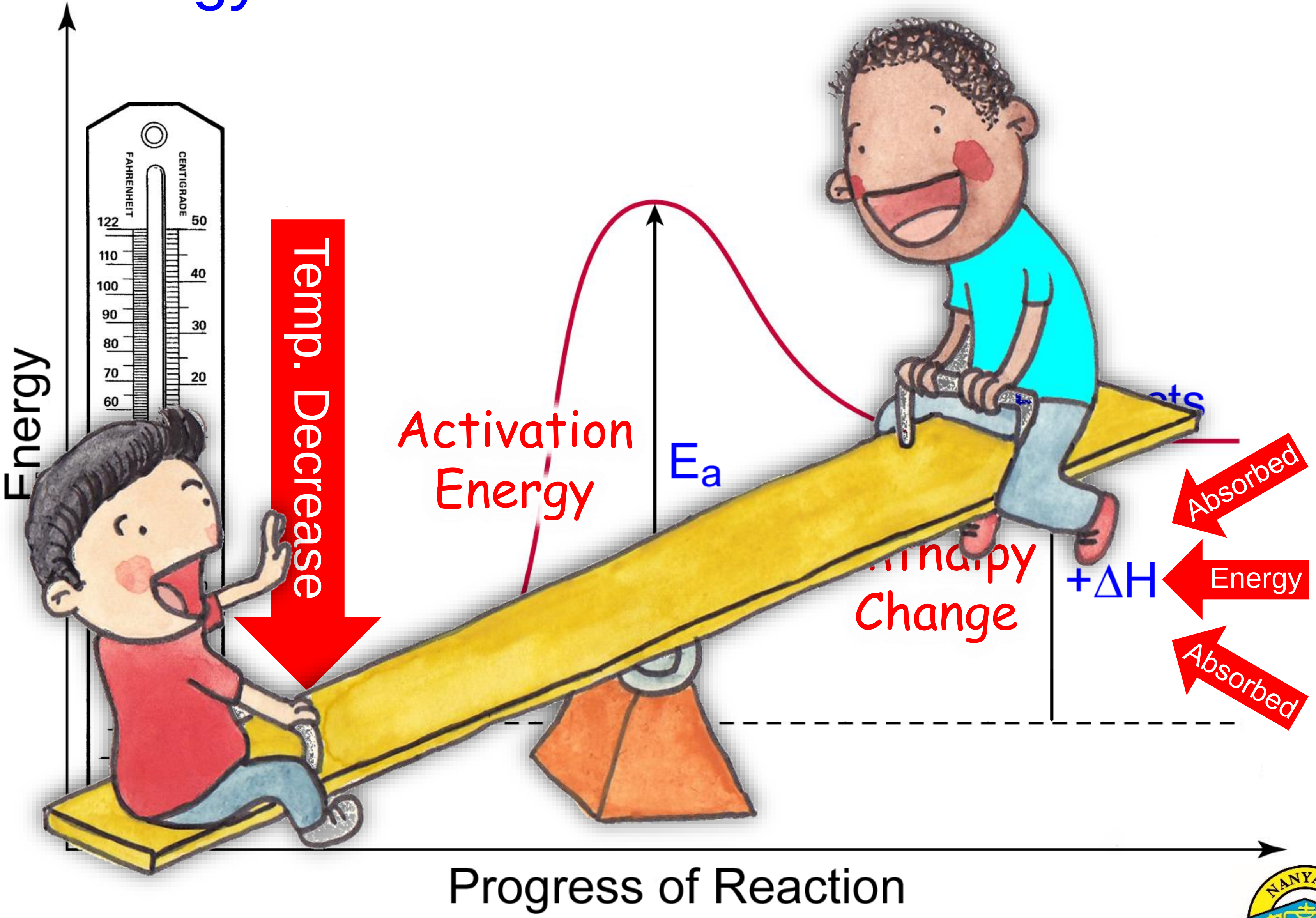
Energy

Absorbed

Energy

Progress of Reaction

Energy Profile for an Endothermic Reaction



Temp. Decrease

Energy

Activation Energy

E_a

ts

nthalpy Change

$+\Delta H$

Absorbed

Energy

Absorbed

Progress of Reaction

Energy Profile for an Endothermic Reaction

What are some examples of *endothermic reactions*?

Energy Profile for an Endothermic Reaction

$$\mathrm{Ba(OH)_2(s) + 2NH_4C\mathit{l}(s) \rightarrow BaC\mathit{l}_2(aq) + 2NH_3(g) + 2H_2O(\mathit{l})}$$

$$\Delta H = +105.1\ \mathrm{kJ\ mol^{-1}}$$

- A few drops of water placed between the flask and wooden block freeze, ‘sticking’ the flask and block together.

Energy Profile for an Endothermic Reaction

Energy Profile for an Endothermic Reaction

- 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

$C_{10}H_{22}(l) \rightarrow C_8H_{18}(l) + C_2H_4(g)$

- Electrolysis:

sodium chloride → sodium + chlorine

$NaCl(l) \rightarrow Na(l) + Cl_2(g)$

Energy from Chemicals – Enthalpy Changes

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 – Enthalpy Changes

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*~~.

Advance Concepts – Entropy

Why do *endothermic* reactions take place? I thought that chemicals reacted in order to *lose* energy!

- In addition to *enthalpy*, we need to consider the system’s *entropy* – how random and chaotic is it?

Advance Concepts – Entropy

- Considering Enthalpy: It is true that chemical reactions take place for higher energy and relatively unstable reactants to form lower energy and relatively stable products. This process is exothermic and will release energy into the surroundings.

- Considering Entropy: It is also true that nature tends to favour systems that are *random and chaotic*. This explains why liquids tend to evaporate – even though the particles are gaining energy, they are also becoming more random and chaotic when they change from a liquid into a gas. *Entropy* is used to describe how random and chaotic a system is – the more chaotic the system, the higher its entropy.

Advance Concepts – Entropy

high energy, ordered reactants → low energy chaotic products

This reaction is very *favourable* and will take place quite easily – it is said to be *spontaneous*.

Advance Concepts – Entropy

high energy, ordered reactants → low energy chaotic products

This reaction is very *favourable* and will take place quite easily – it is said to be *spontaneous*.

low energy chaotic reactants → high energy, ordered products

This reaction is *unfavourable* and will take place with difficulty – it is said to be *non-spontaneous*.

Advance Concepts – Entropy

high energy, ordered reactants → low energy chaotic products

This reaction is very *favourable* and will take place quite easily – it is said to be *spontaneous*.

low energy chaotic reactants → high energy, ordered products

This reaction is *unfavourable* and will take place with difficulty – it is said to be *non-spontaneous*.

low energy ordered reactants → high energy, chaotic products

Although this reaction is *unfavourable* due to the increase in energy, it could still be *spontaneous* if there is a large *increase in entropy* (*i.e.* there is a relatively *small* increase in *energy*, but the system becomes *very chaotic*). This change is *endothermic*.

Advance Concepts – Entropy

- *Gibbs free energy* gives Chemists an idea as to whether a chemical change will be either *spontaneous* or *non-spontaneous*.

- Gibbs free energy is defined by the equation:

$$\Delta G = \Delta H - T\Delta S$$

Where:

ΔG = change in Gibbs free energy / J

ΔH = enthalpy change of the system / J

T = temperature of the system / K

ΔS = entropy change of the system / $J \cdot K^{-1}$

- If ΔG is *negative*, then the chemical change is *spontaneous*.

Advance Concepts – Entropy

- Gibbs free energy is defined by the equation:

$$\Delta G = \Delta H - T\Delta S$$

	$\Delta H < 0$	$\Delta H > 0$
$\Delta S > 0$	spontaneous at all temperatures ($\Delta G < 0$)	spontaneous at high temperatures (when $T\Delta S$ is large)
$\Delta S < 0$	spontaneous at low temperatures (when $T\Delta S$ is small)	non-spontaneous at all temperatures ($\Delta G > 0$)

Energy from Chemicals – Enthalpy Changes

Which reaction is more common, *exothermic* or *endothermic*?

Energy from Chemicals – Enthalpy Changes

- 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.

Energy from Chemicals – Enthalpy Changes

- 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.

Energy from Chemicals – Enthalpy Changes

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

- Exothermic reactions are therefore more common than endothermic reactions.

Energy from Chemicals – Enthalpy Changes

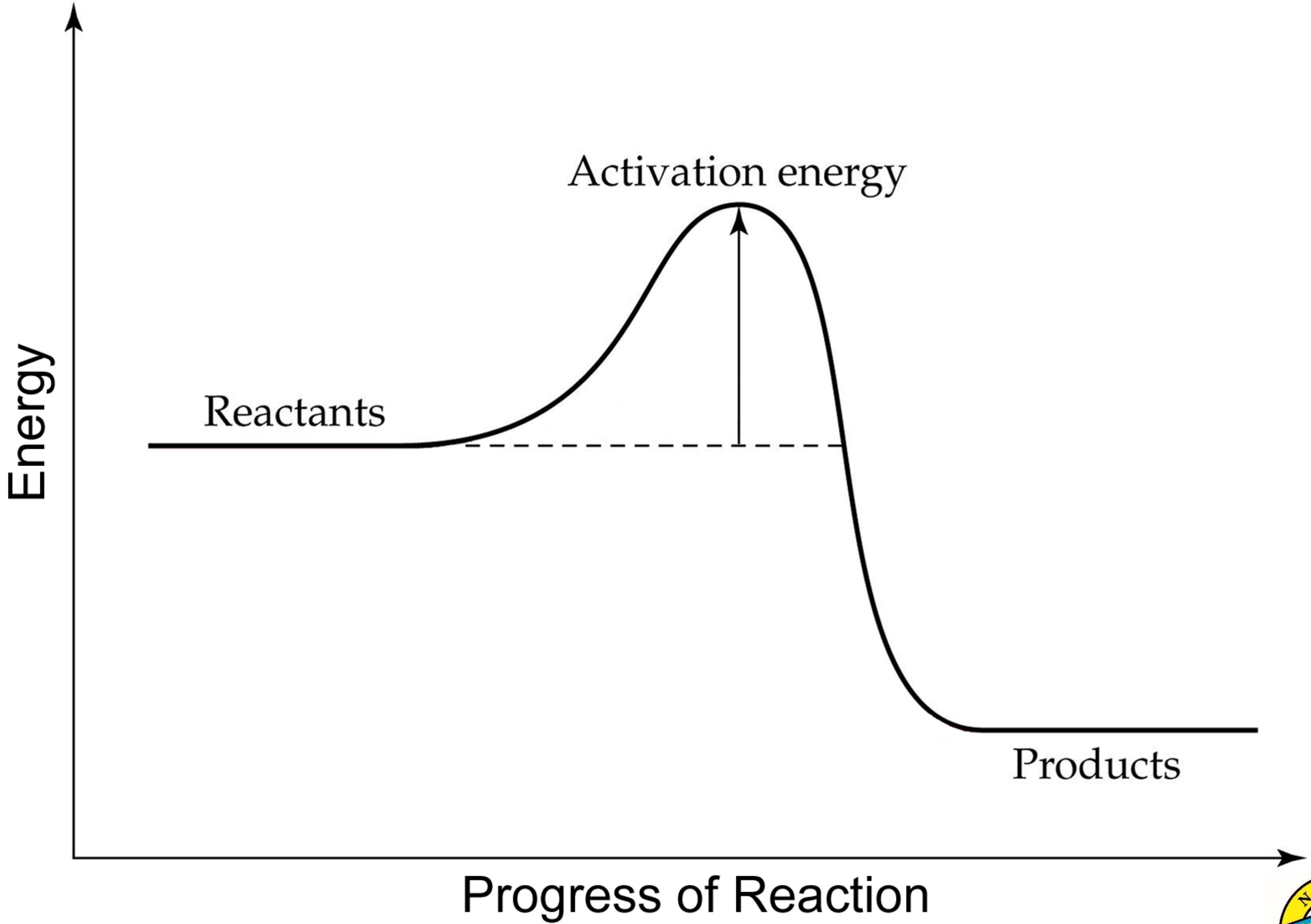
- Question: How do these diagrams illustrate the concept of activation energy?

- 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.

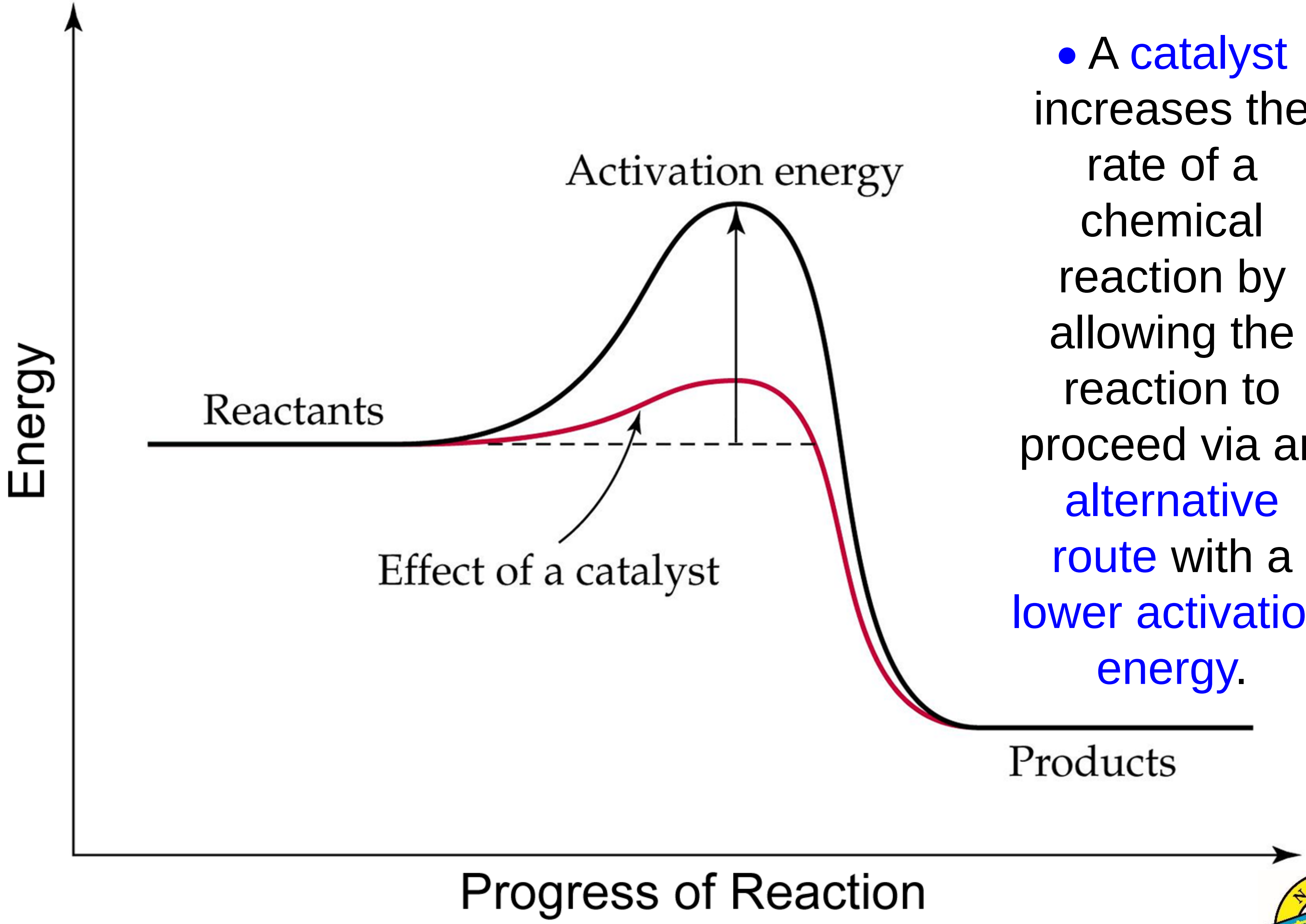
Effect of a Catalyst on Activation Energy

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

Effect of a Catalyst on Activation Energy

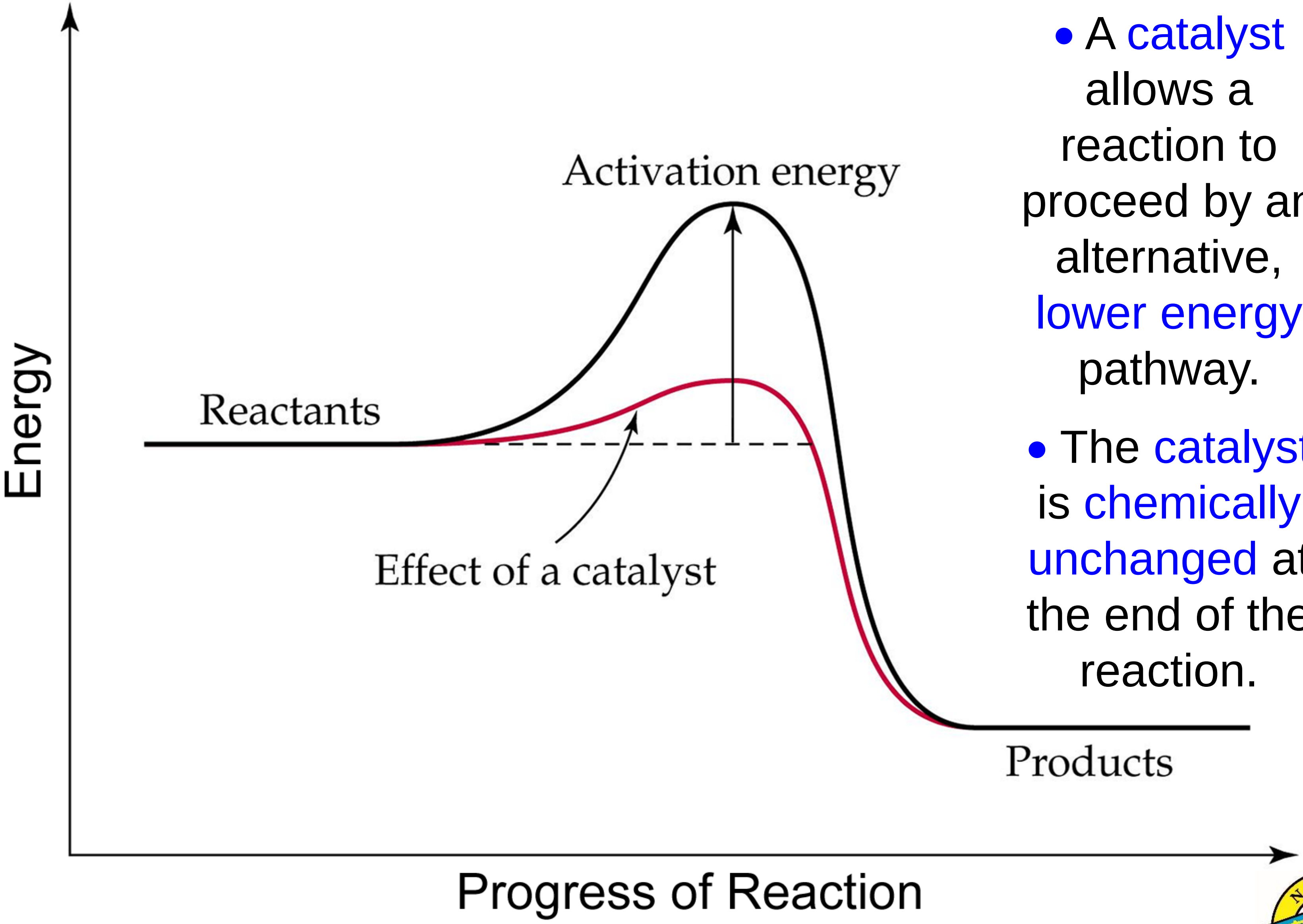


Effect of a Catalyst on Activation Energy



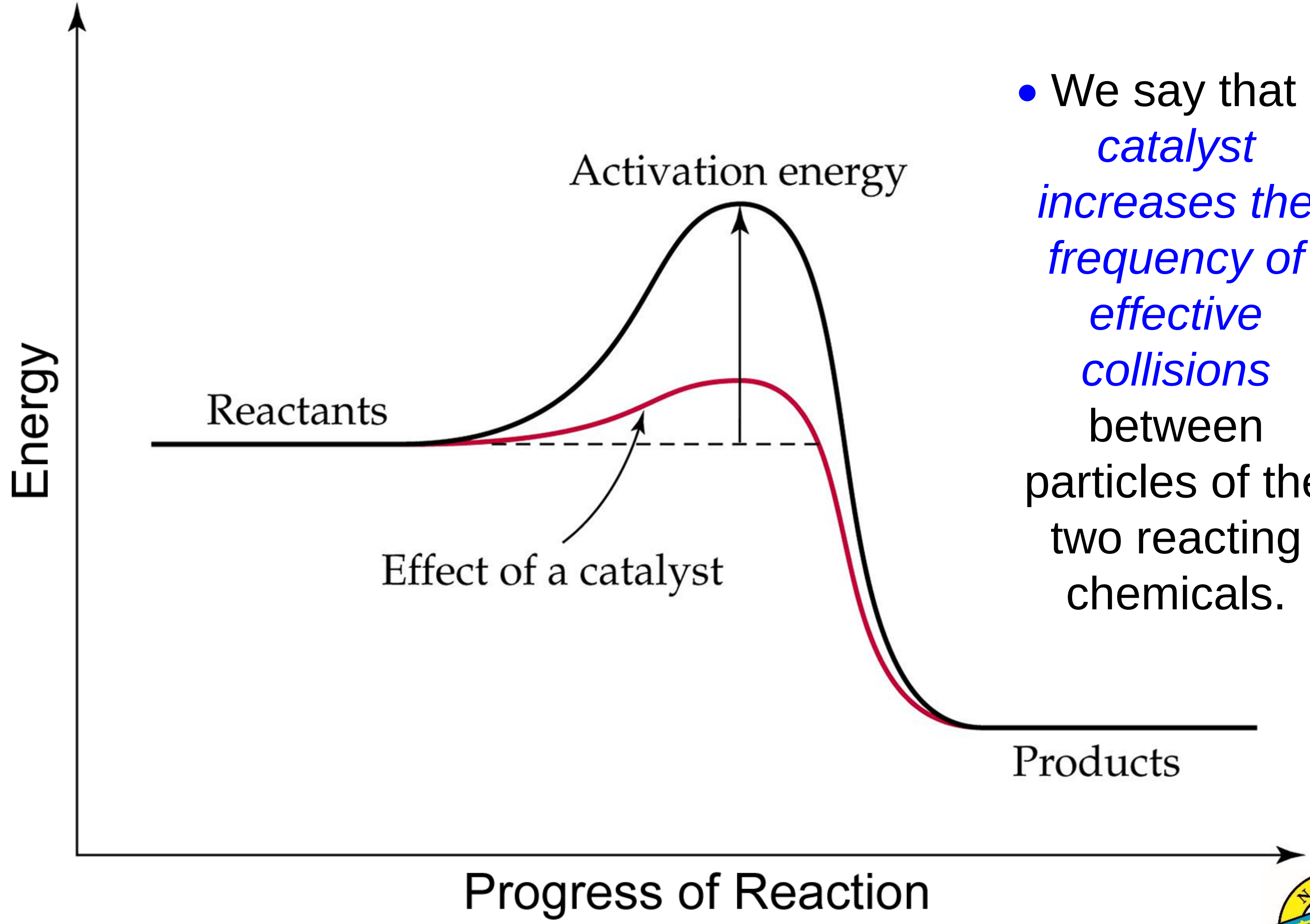
- A catalyst increases the rate of a chemical reaction by allowing the reaction to proceed via an alternative route with a lower activation energy.

Effect of a Catalyst on Activation Energy



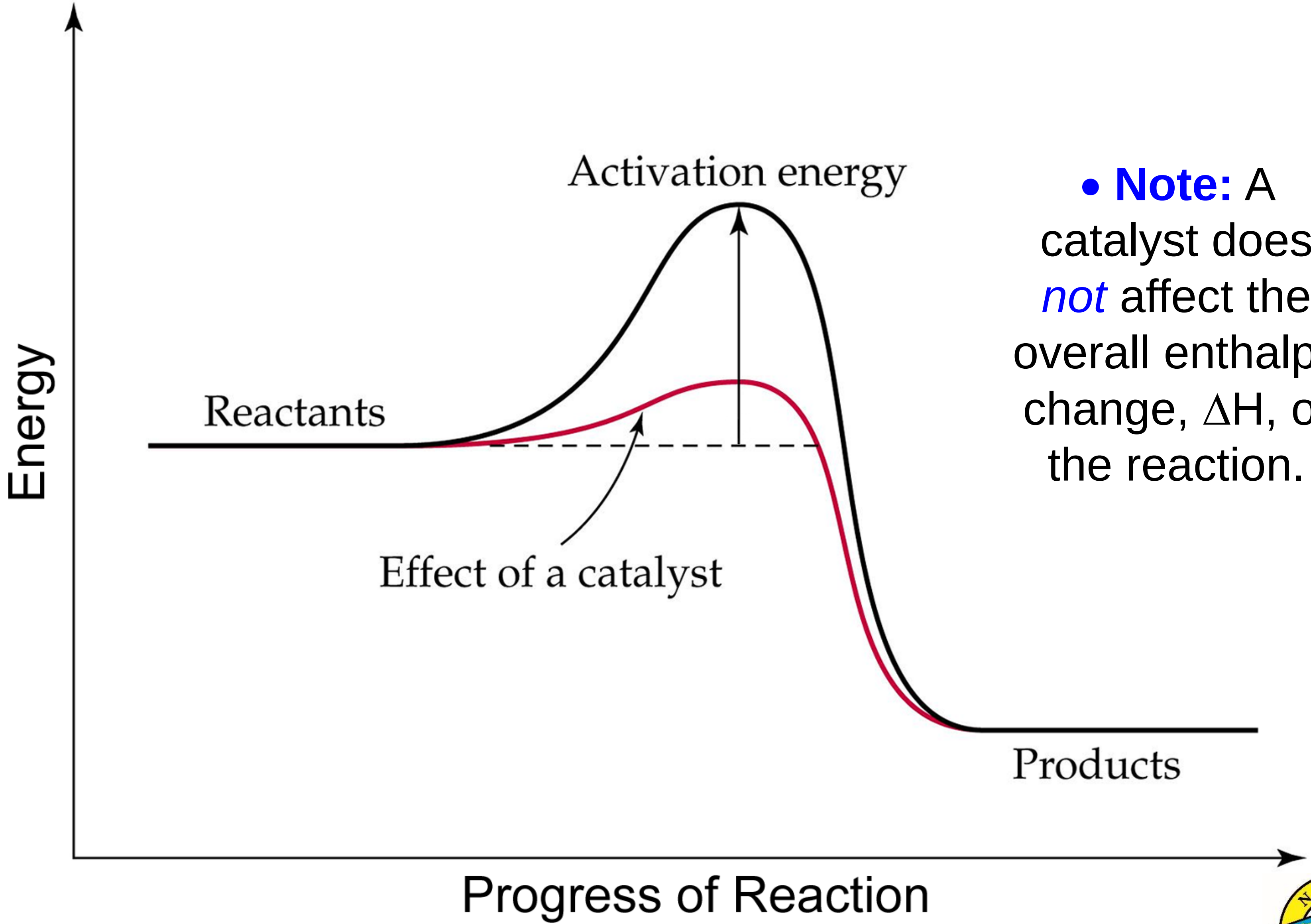
- A catalyst allows a reaction to proceed by an alternative, lower energy pathway.
- The catalyst is chemically unchanged at the end of the reaction.

Effect of a Catalyst on Activation Energy



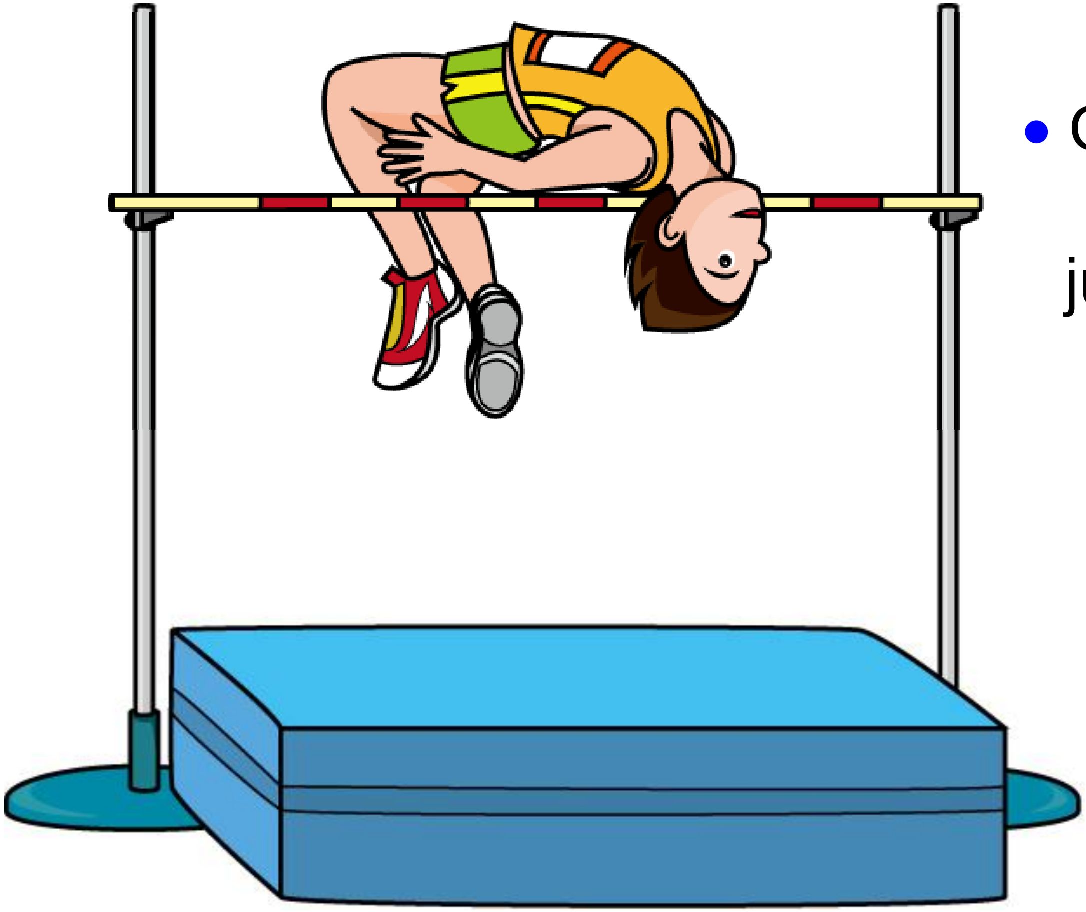
- We say that a *catalyst increases the frequency of effective collisions* between particles of the two reacting chemicals.

Effect of a Catalyst on Activation Energy



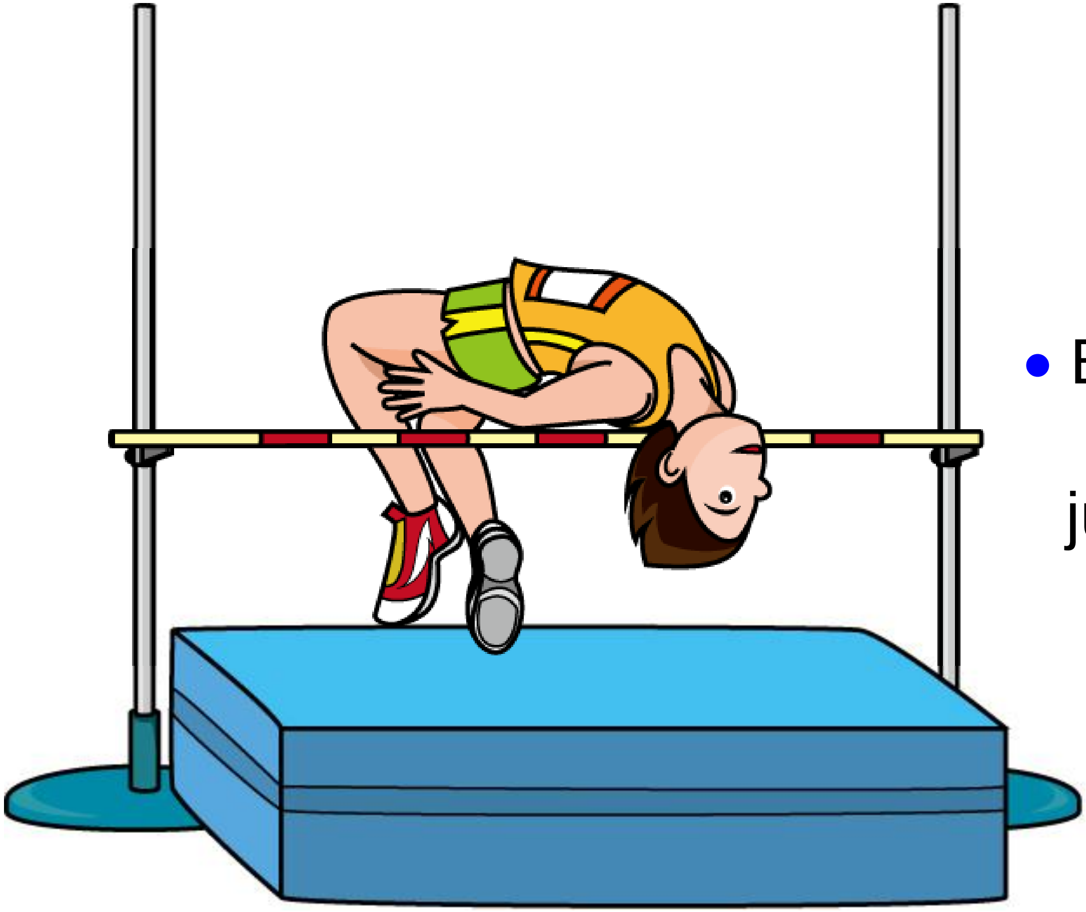
- **Note:** A catalyst does *not* affect the overall enthalpy change, ΔH, of the reaction.

Effect of a Catalyst on Activation Energy



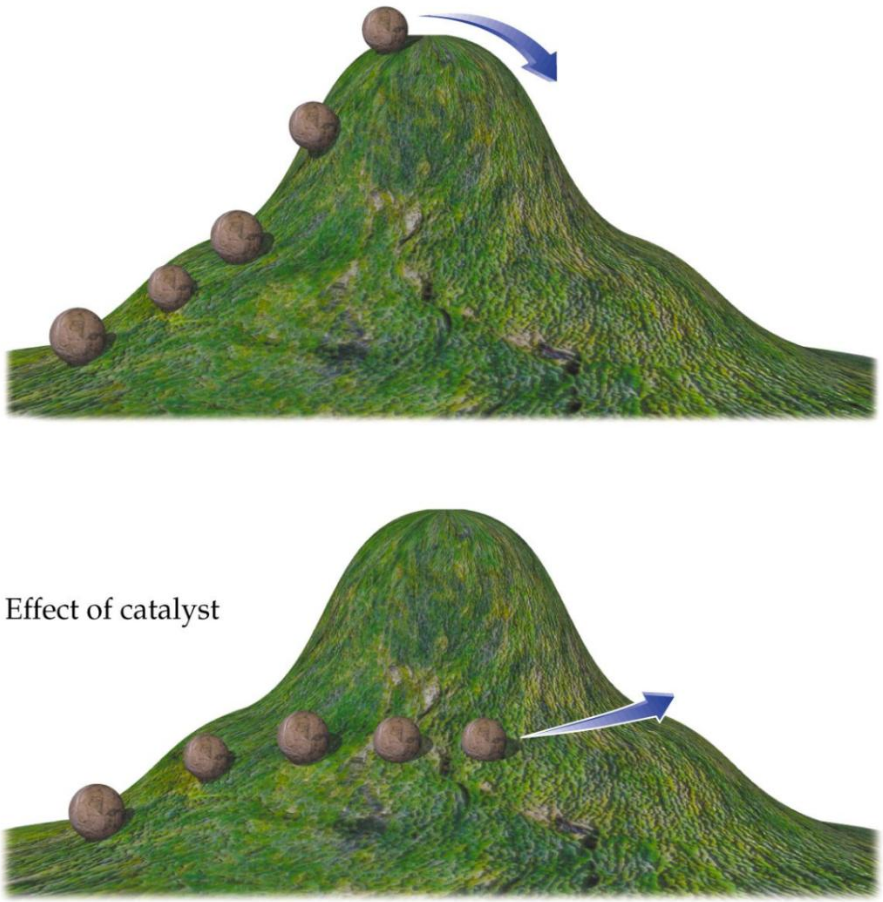
- Only 3 out of 30 students can jump this high.

Effect of a Catalyst on Activation Energy



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

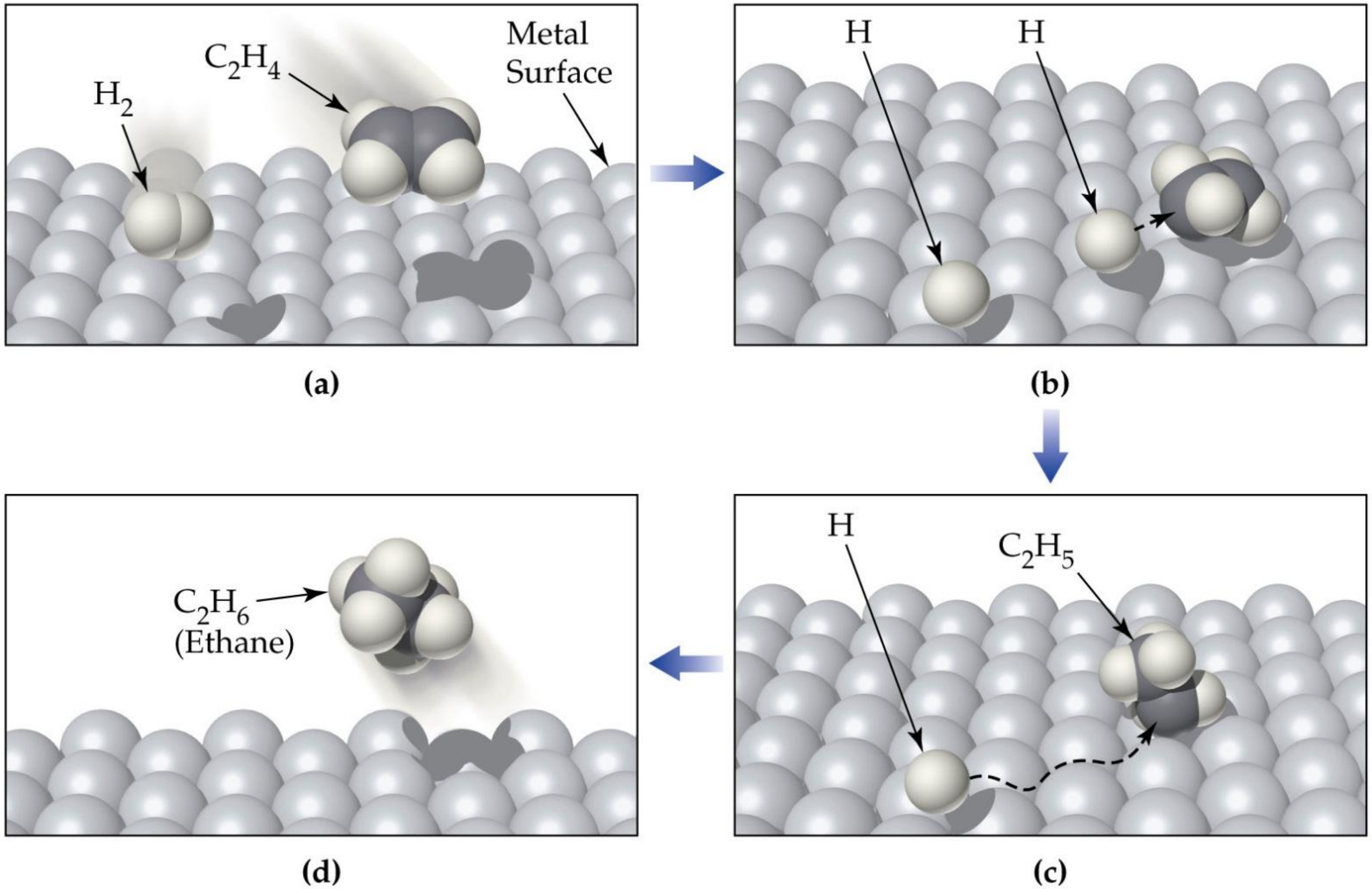
Effect of a Catalyst on Activation Energy

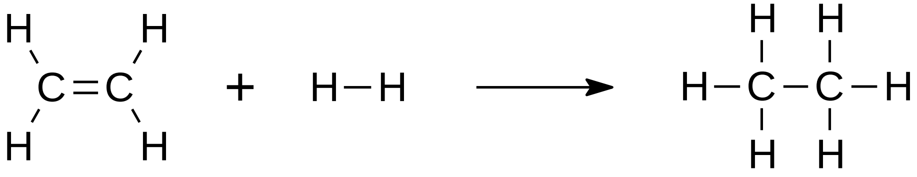


- Analogy for activation energy. Less energy is required to push the boulders around the side of the hill compared to the energy that is required to push the boulders over the hill.

- 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 on Activation Energy





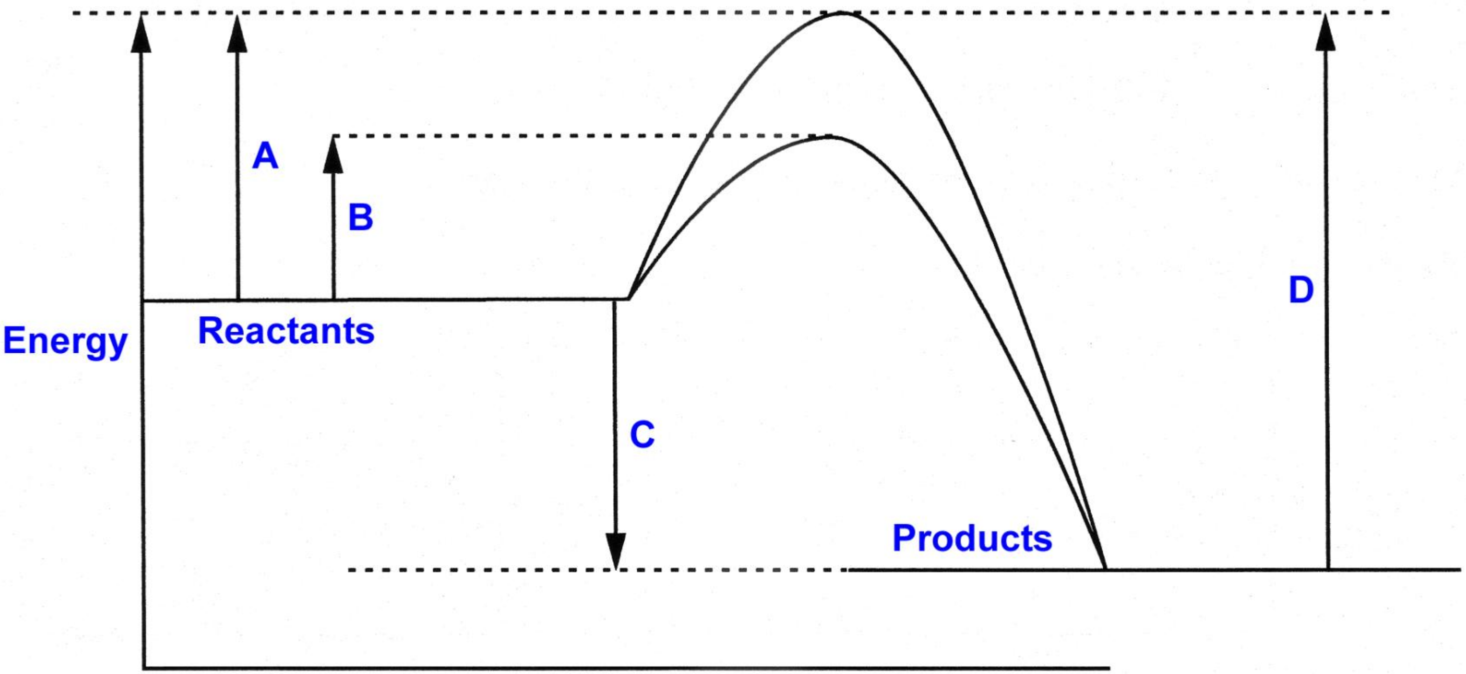
Ethene Hydrogen Ethane

Assessment for Learning

Can I please have a short question to check my current understanding?

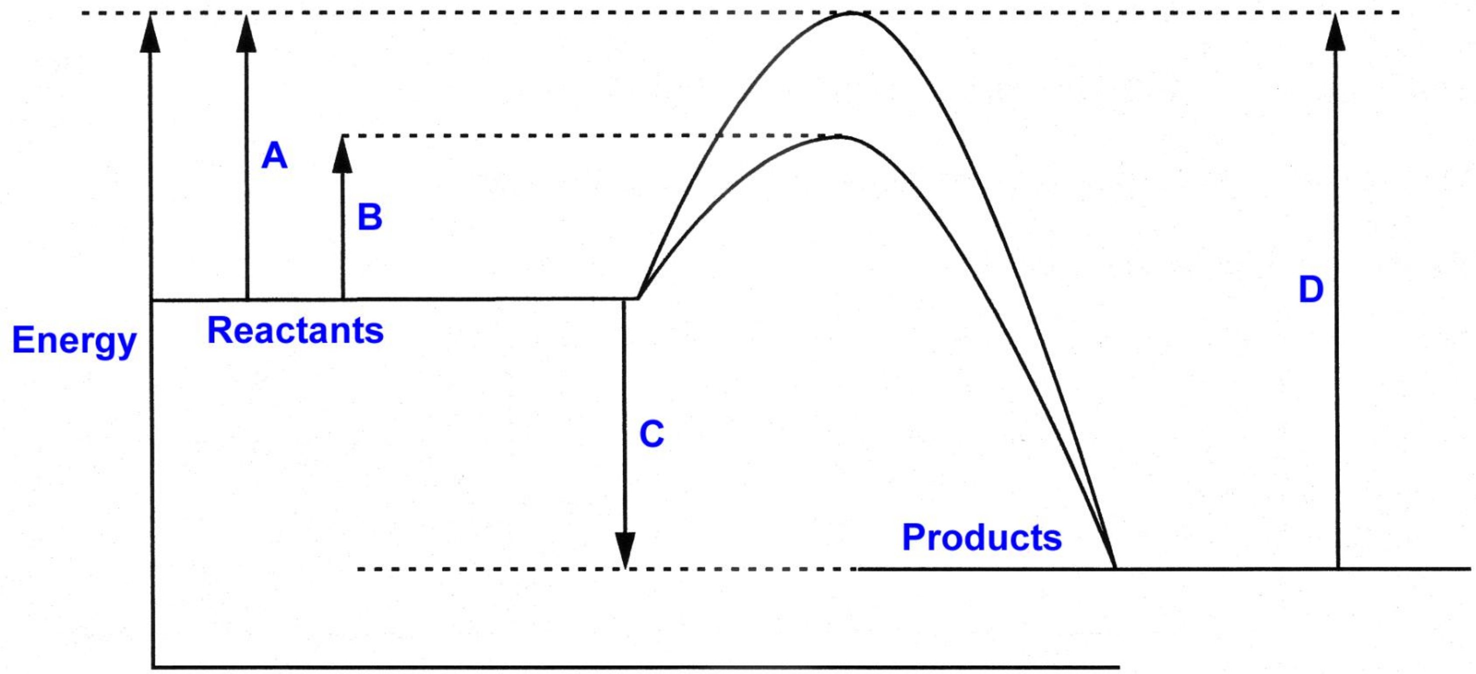
Assessment for Learning

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



Assessment for Learning

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



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.

Enthalpy Change Calculations

How can I *calculate* the *enthalpy change* of a chemical reaction?

Enthalpy Change Calculations

- 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*".

Enthalpy Change Calculations

- 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*".

Enthalpy 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.

Enthalpy Change Calculations

Important Things to Note…

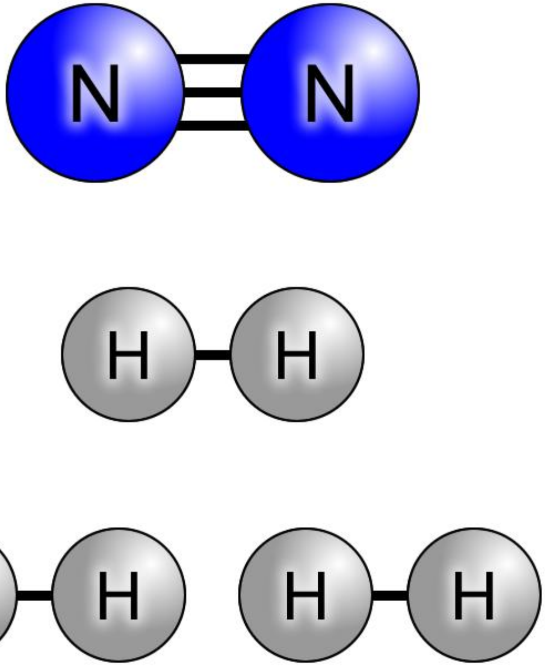
- Bond breaking *absorbs energy* and is therefore endothermic. Numerical values of ΔH for bond breaking are positive.

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

Enthalpy Change Calculations

nitrogen + hydrogen ⇌ ammonia

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



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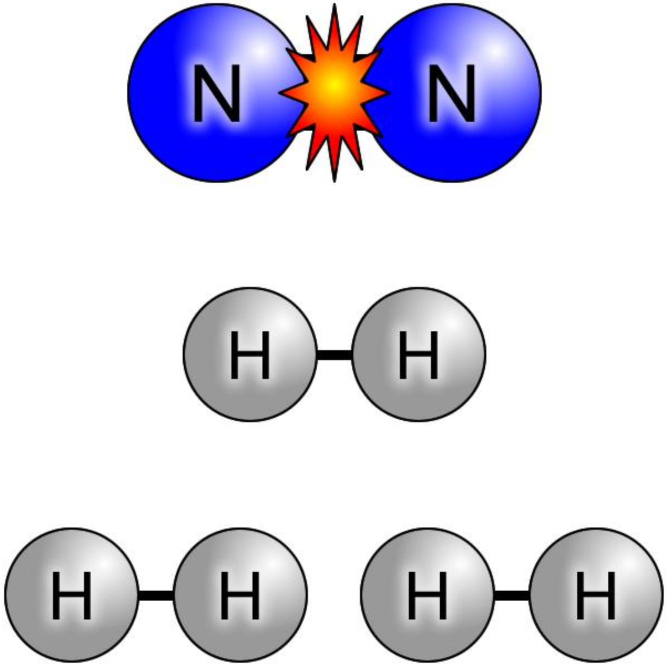
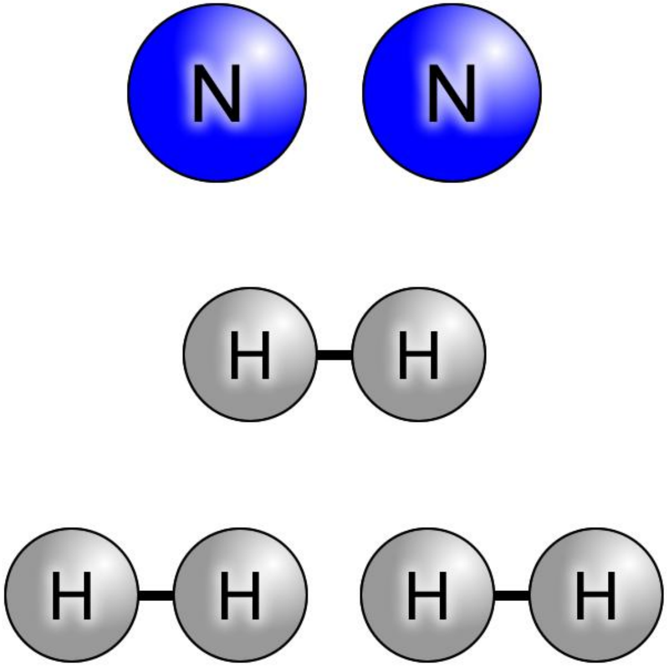


1) Bond breaking *absorbs* energy. *ΔH is positive*.

Enthalpy Change Calculations

nitrogen + hydrogen ⇌ ammonia

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



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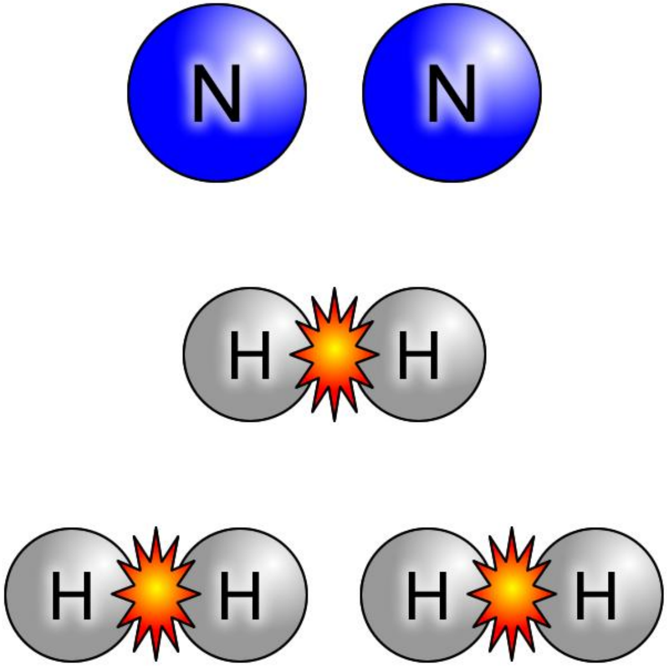

1) Bond breaking *absorbs* energy.

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Enthalpy Change Calculations

nitrogen + hydrogen ⇌ ammonia

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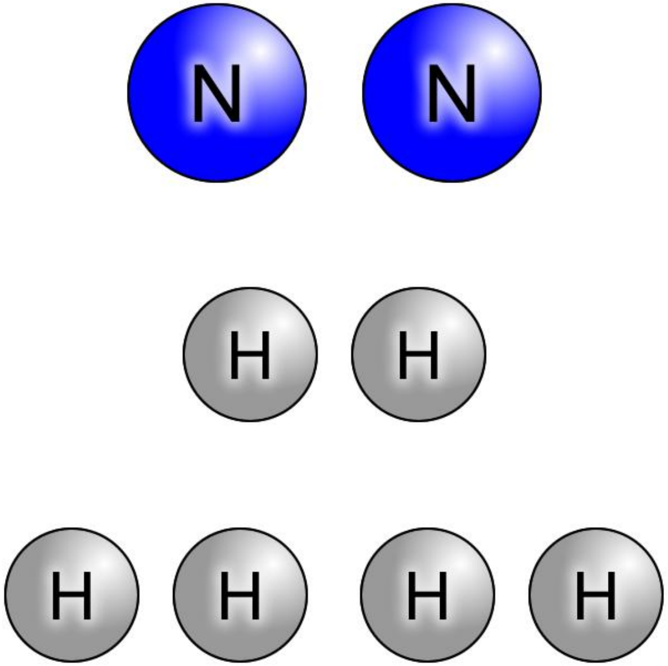

Enthalpy Change Calculations

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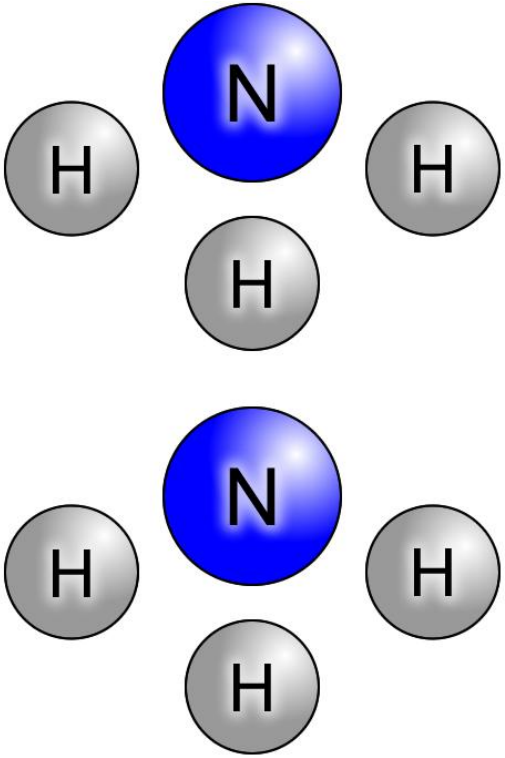
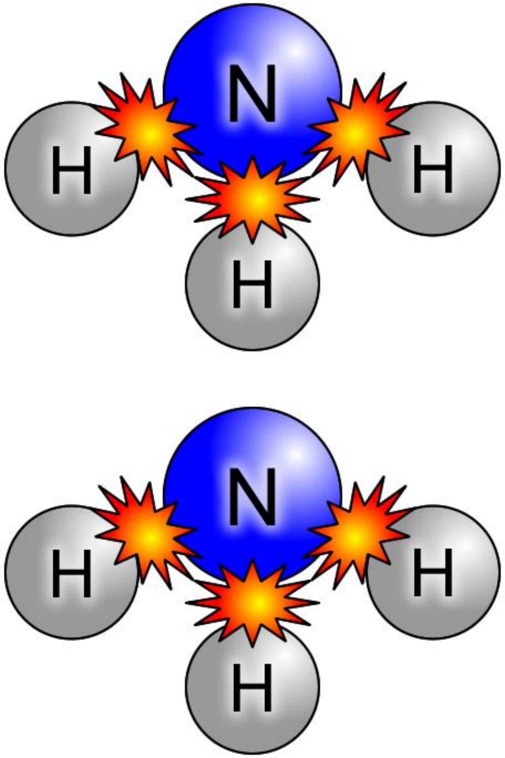


Enthalpy Change Calculations

nitrogen + hydrogen $\rightleftharpoons$ ammonia

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

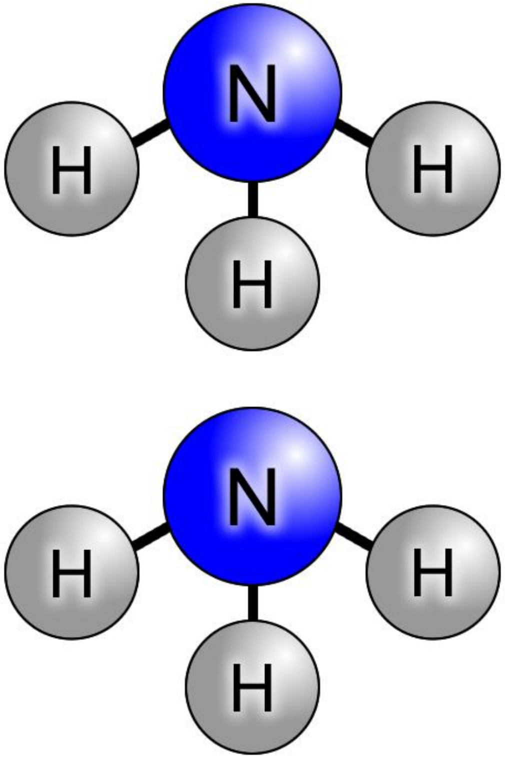


2) Bond formation *releases* energy.
ΔH is negative.

Enthalpy Change Calculations

nitrogen + hydrogen ⇌ ammonia

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



Enthalpy Change Calculations

Important Things to Note…

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

Enthalpy 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.

Enthalpy 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*.

Enthalpy 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.

Enthalpy Change Calculations

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

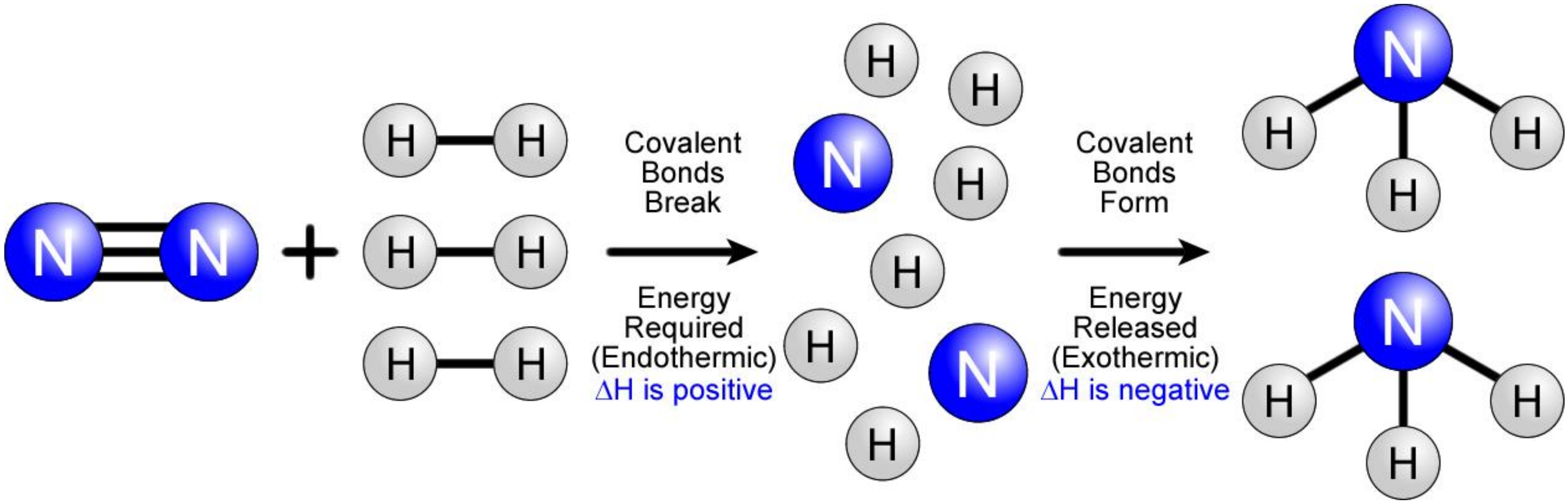
- Overall, the reaction *releases energy* into the surroundings. The reaction is therefore *exothermic*.

Enthalpy 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:



Covalent Bonds Break

Energy Required (Endothermic)
ΔH is positive

Covalent Bonds Form

Energy Released (Exothermic)
ΔH is negative

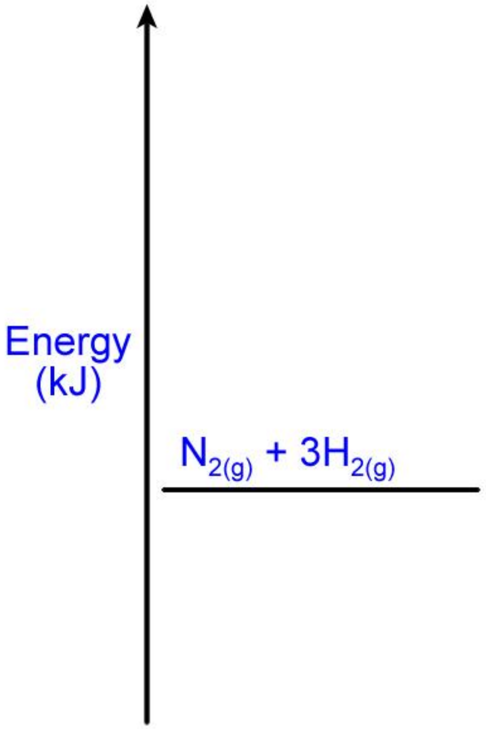
Enthalpy Change Calculations – Example #1:

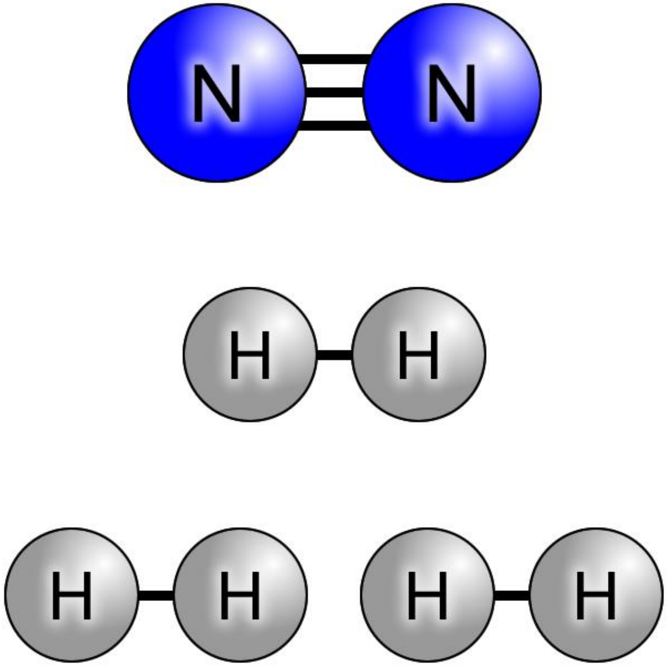
The Reaction Between Nitrogen and Hydrogen to form Ammonia

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

$N{\equiv}N = 944\ kJ\ mol^{-1}$ $H{-}H = 436\ kJ\ mol^{-1}$ $N{-}H = 388\ kJ\ mol^{-1}$

energy level diagram





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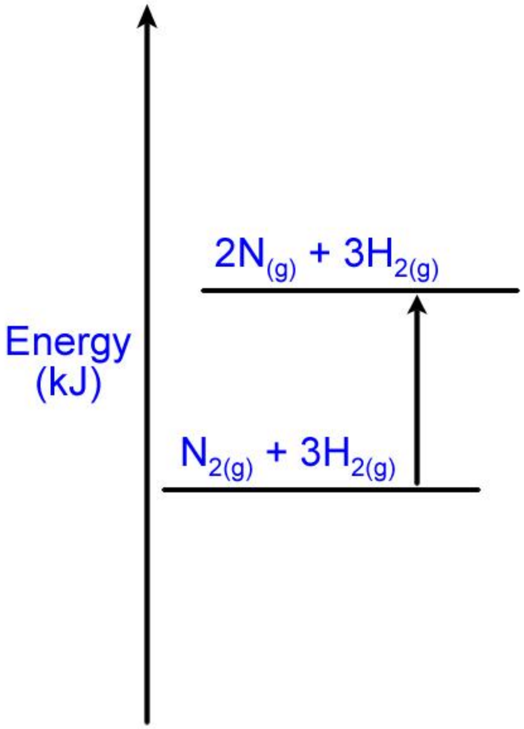

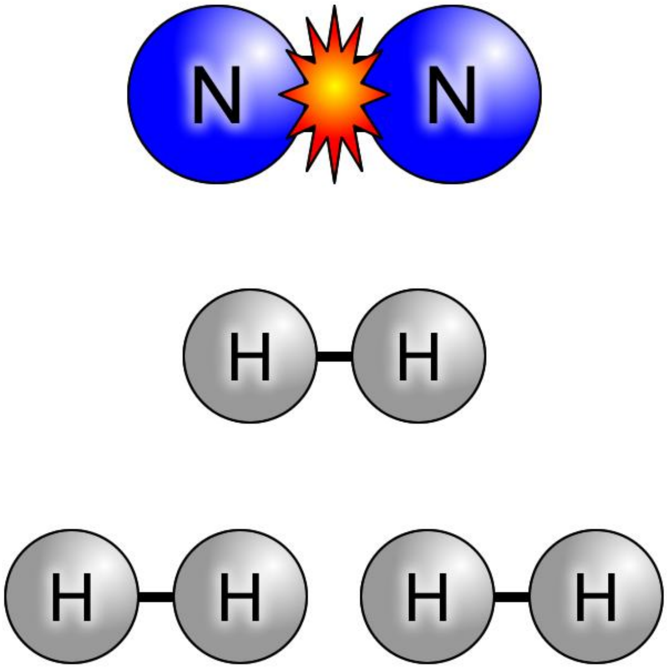

1 × N≡N bond is broken forming 2 × N atoms.

Enthalpy Change Calculations – Example #1:

The Reaction Between Nitrogen and Hydrogen to form Ammonia

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

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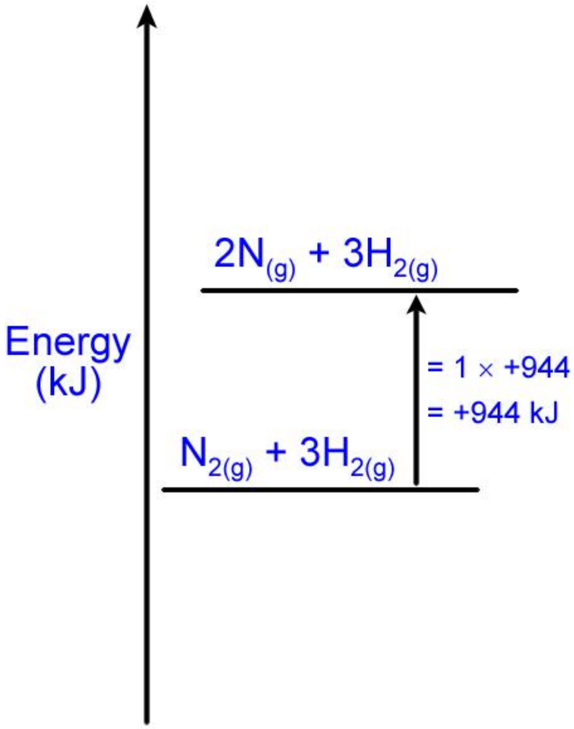

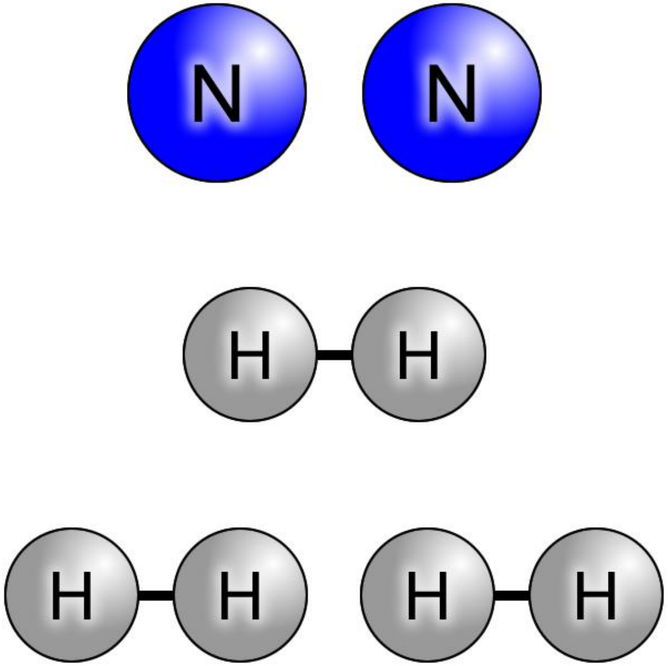

Bond breaking is *endothermic*: ΔH for this change is *positive*.

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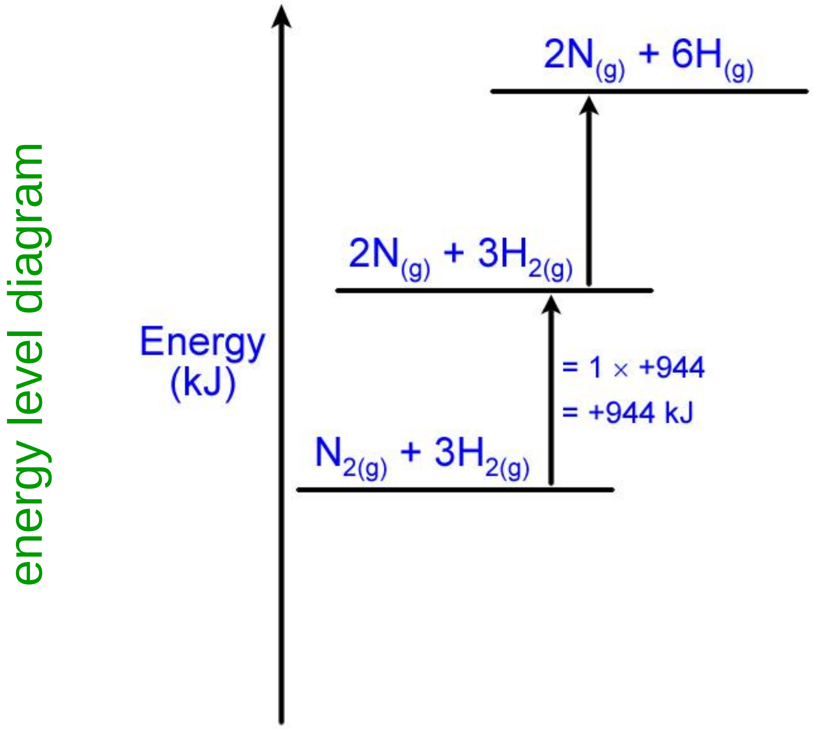

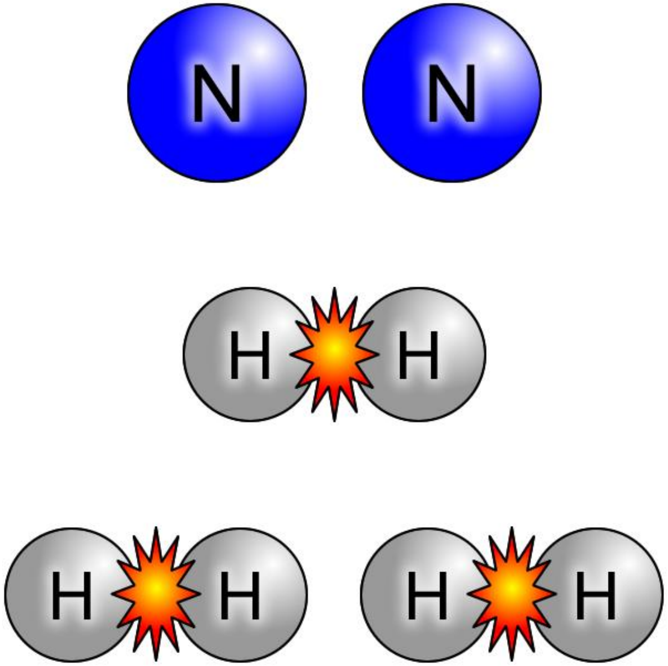

3 × H–H bonds are broken forming 6 × H atoms.

Enthalpy Change Calculations – Example #1:

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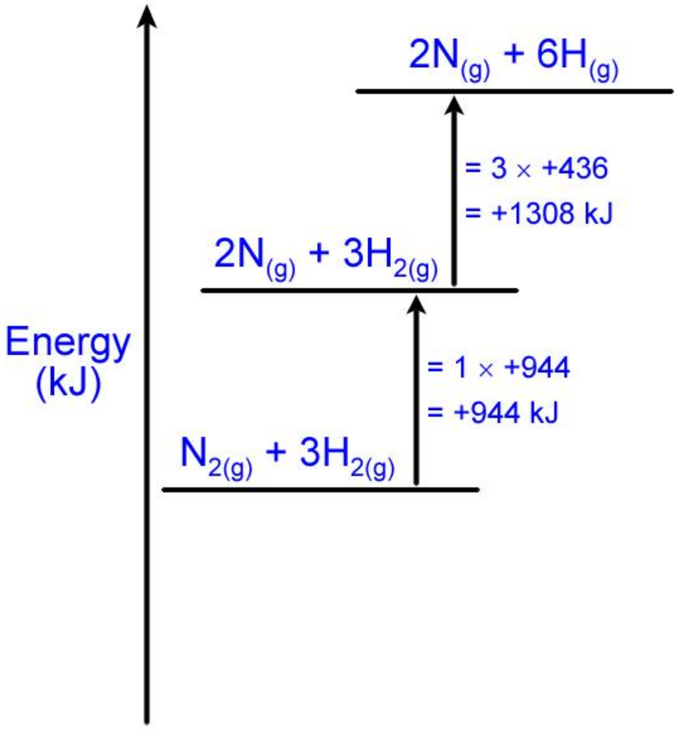

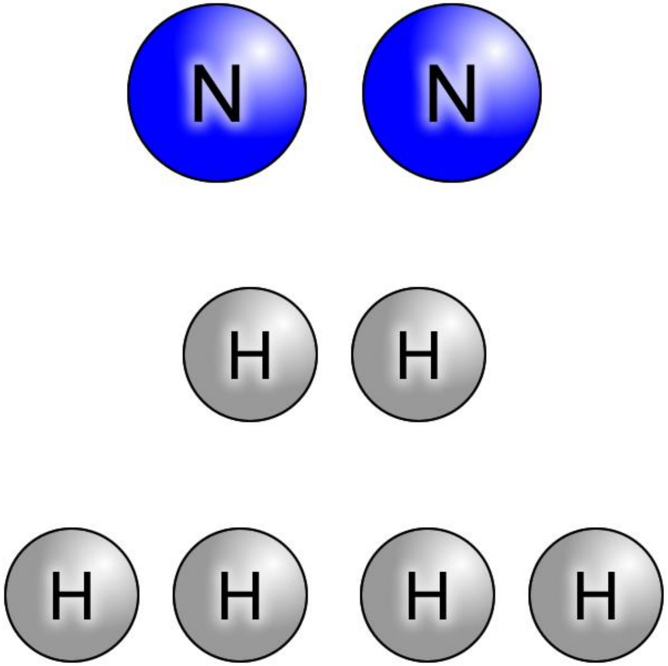

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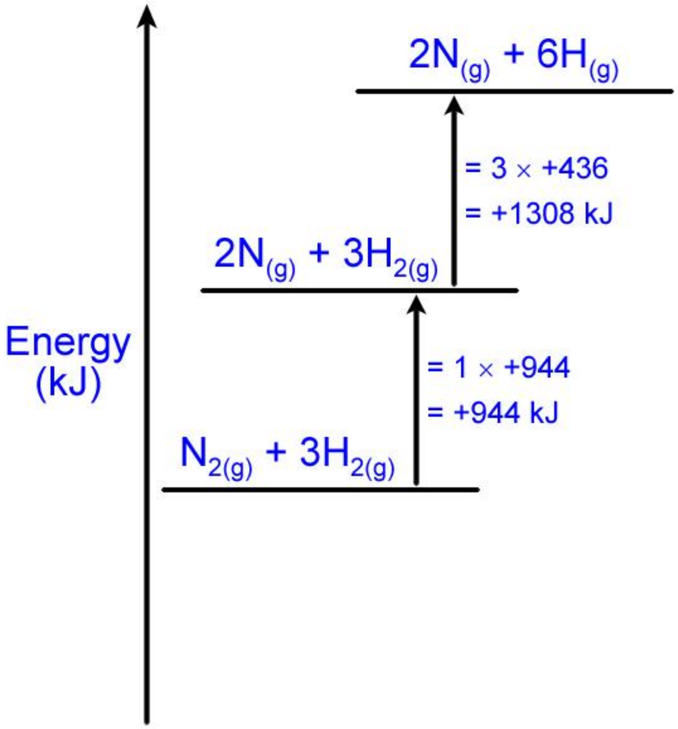

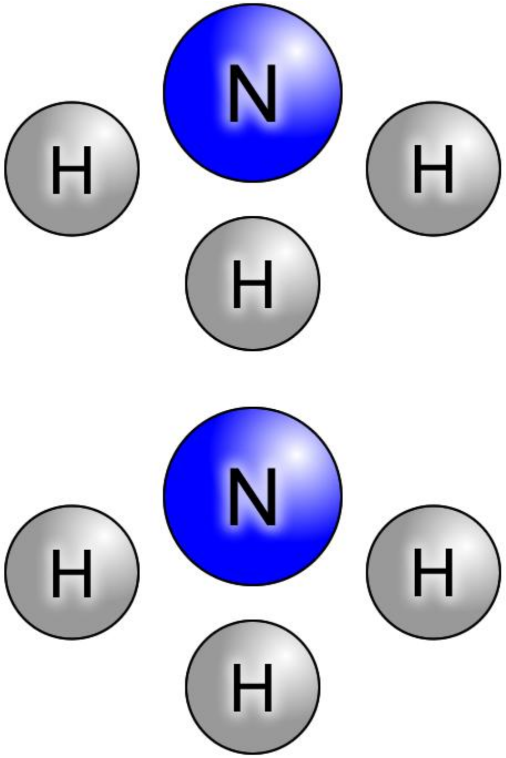

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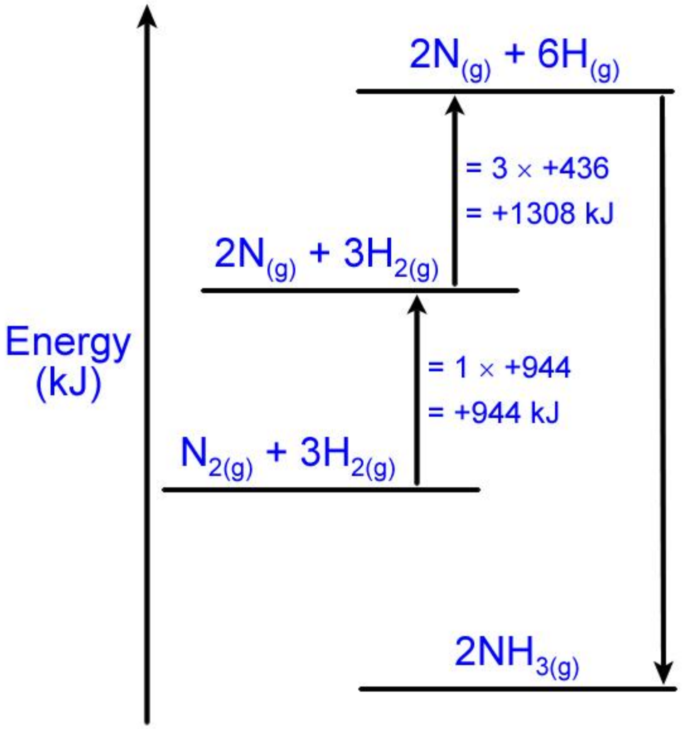

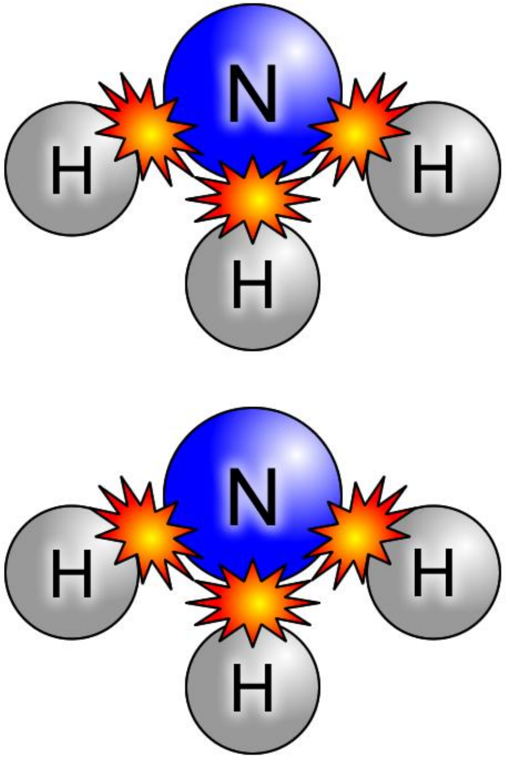

6 × N–H bonds are formed creating 2 × NH_3 molecules.

Enthalpy Change Calculations – Example #1:

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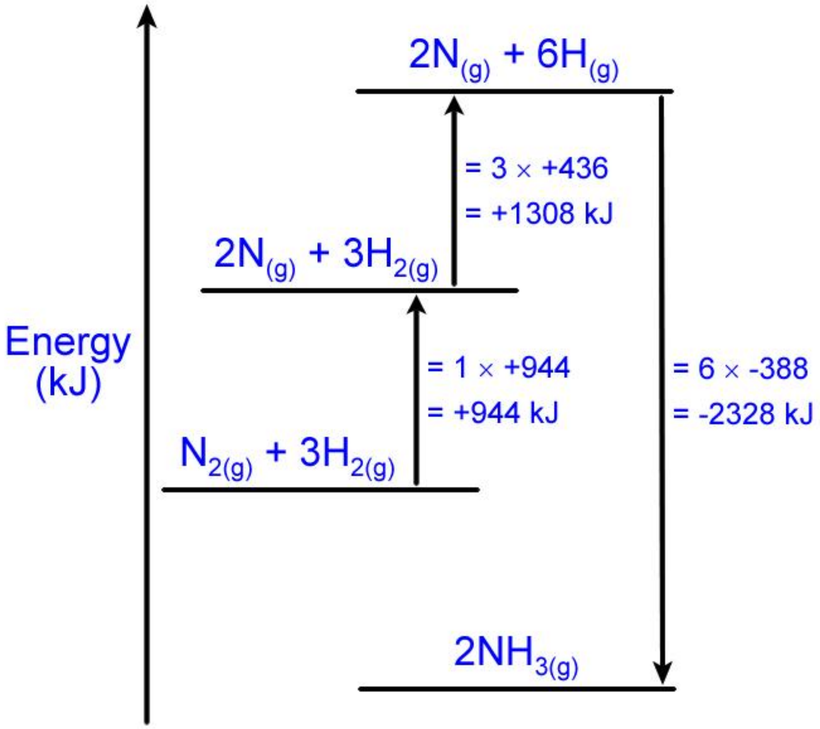

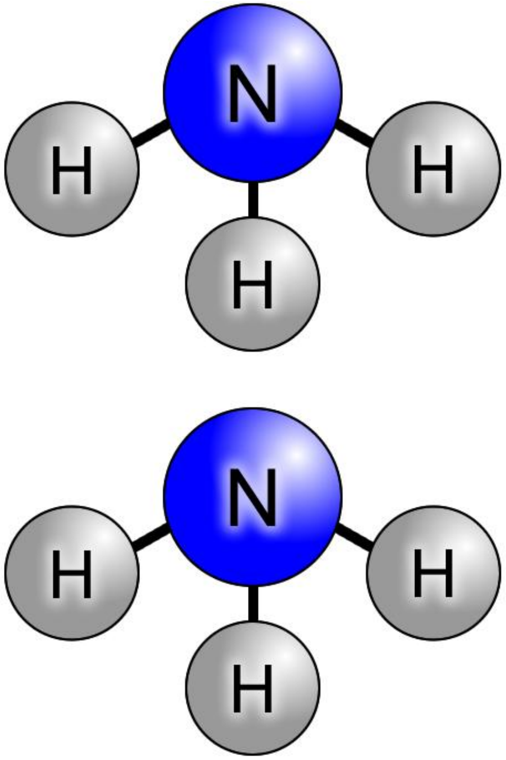

Bond formation is *exothermic*: ΔH for this change is *negative*.

Enthalpy Change Calculations – Example #1:

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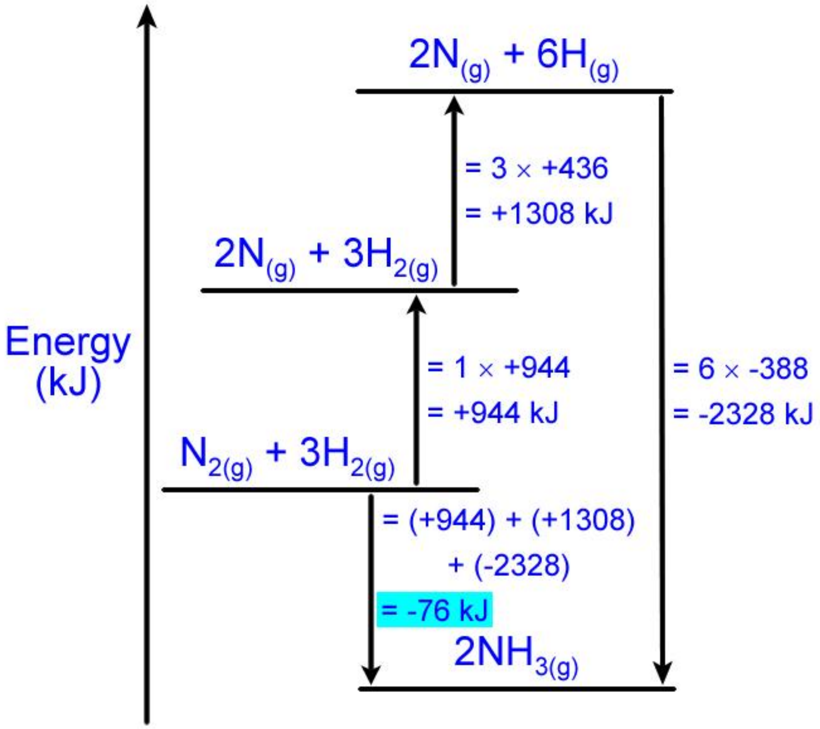

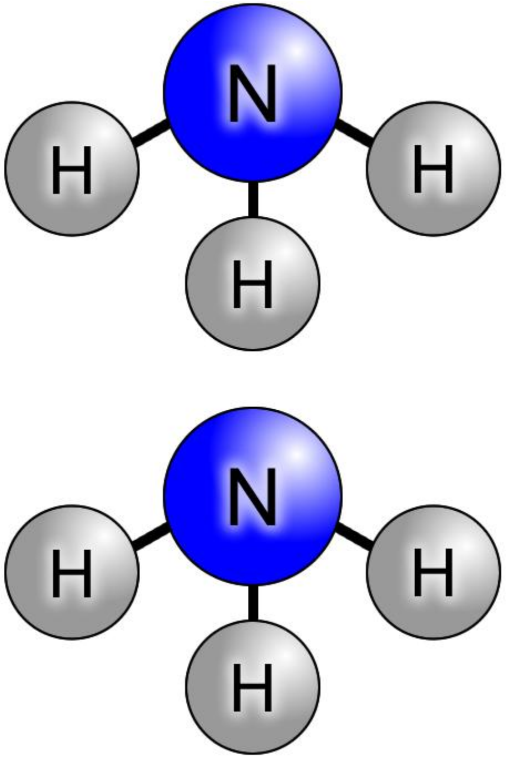

$$(+944) + (+1308) + (-2328) = -76 \text{ kJ mol}^{-1}$$

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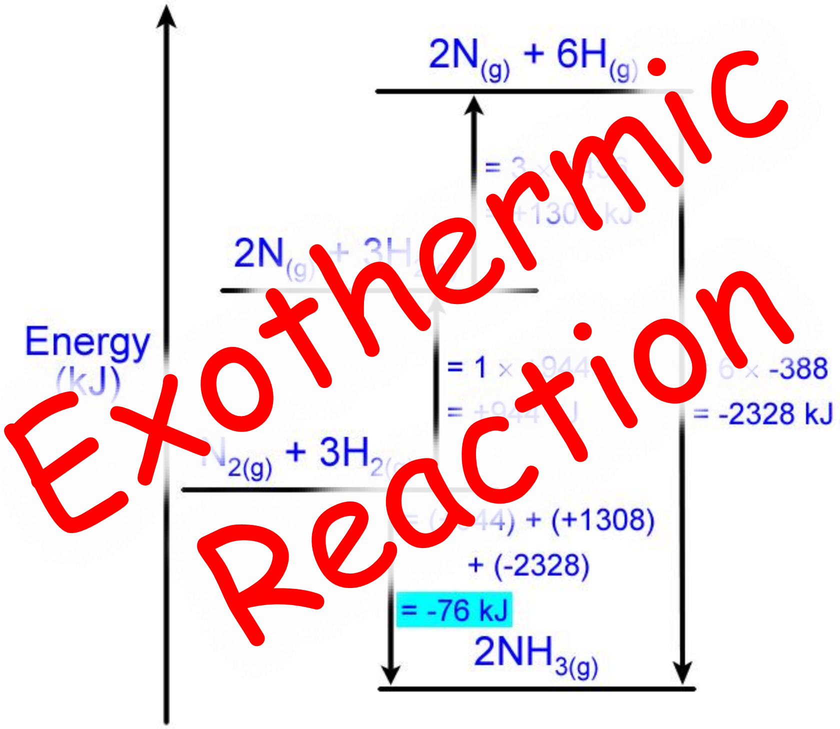
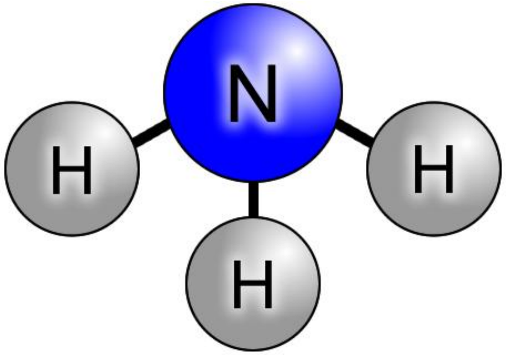
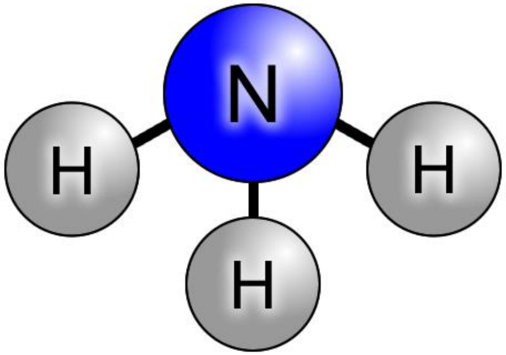






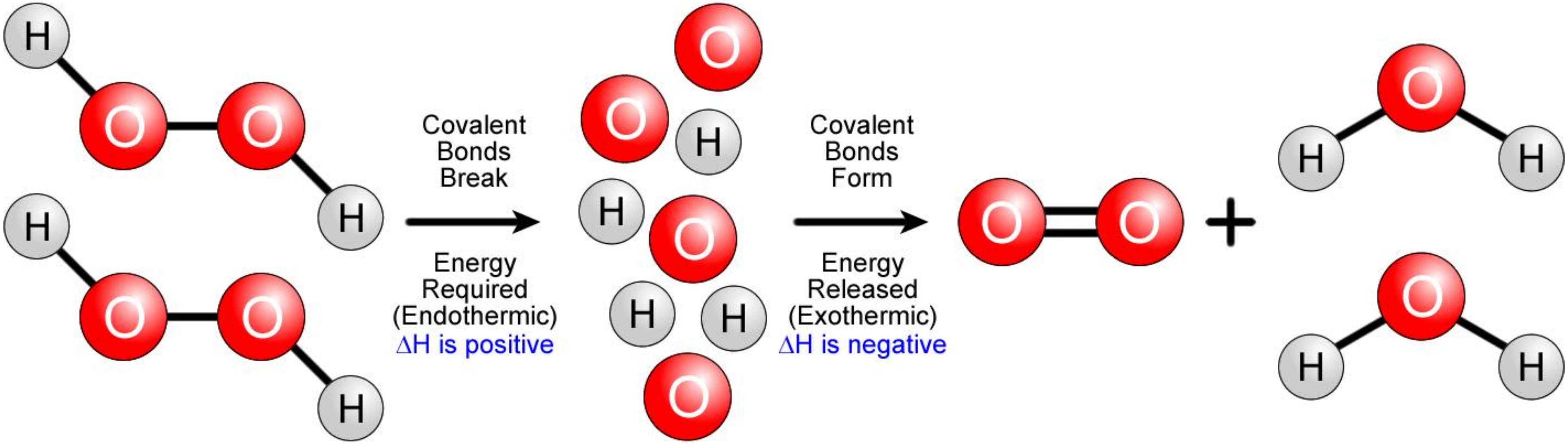
$$(+944) + (+1308) + (-2328) = -76\ kJ\ mol^{-1}$$

Enthalpy 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:



Covalent Bonds Break

Energy Required (Endothermic)

ΔH is positive

Covalent Bonds Form

Energy Released (Exothermic)

ΔH is negative

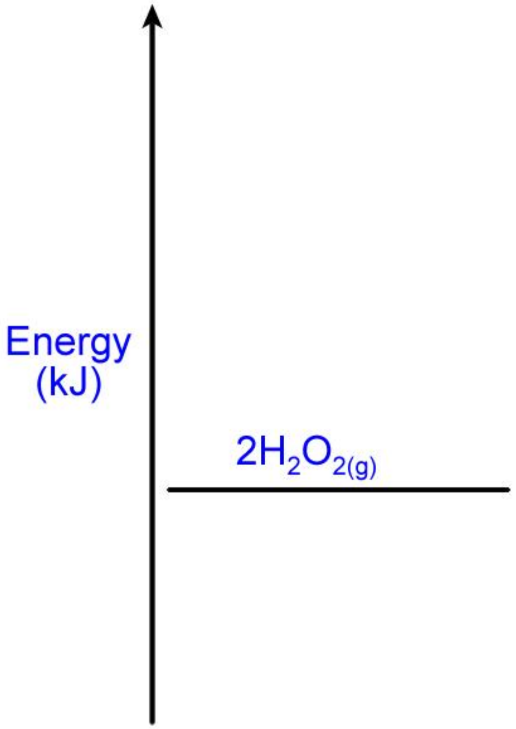
Enthalpy Change Calculations – Example #2:

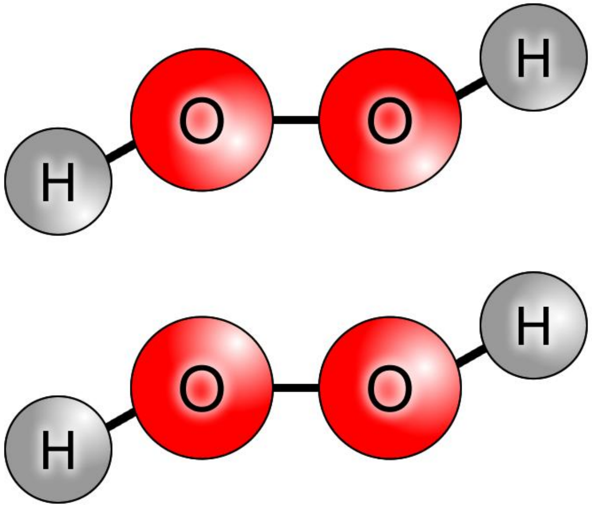
The Decomposition of Hydrogen Peroxide to form Oxygen and Water

$$2H_2O_2(g) \rightarrow O_2(g) + 2H_2O(g)$$

O–O = 146 kJ mol^{-1} O–H = 463 kJ mol^{-1} O=O = 496 kJ mol^{-1}

energy level diagram





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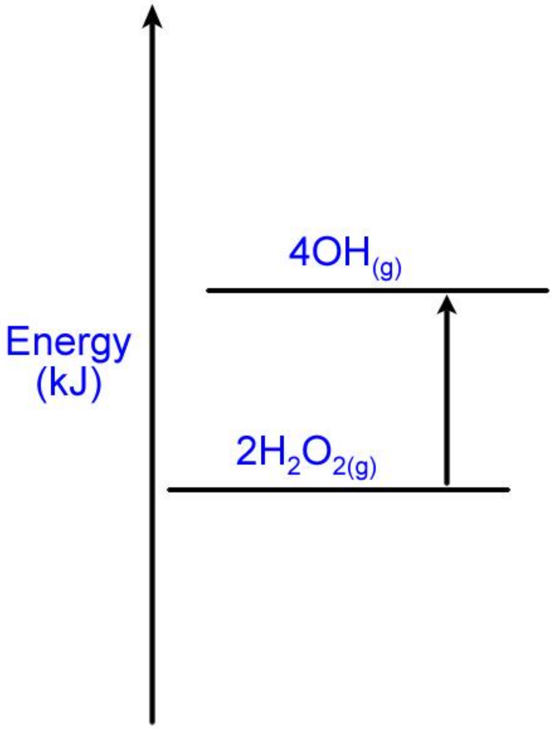

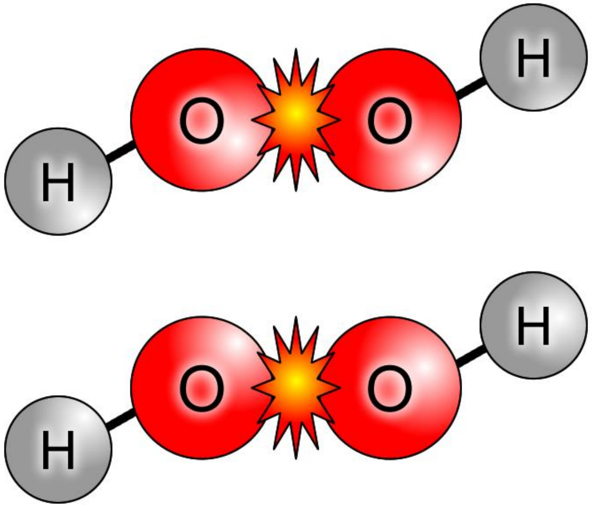

2 × O–O bond are broken forming 4 × OH molecules.

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The Decomposition of Hydrogen Peroxide to form Oxygen and Water

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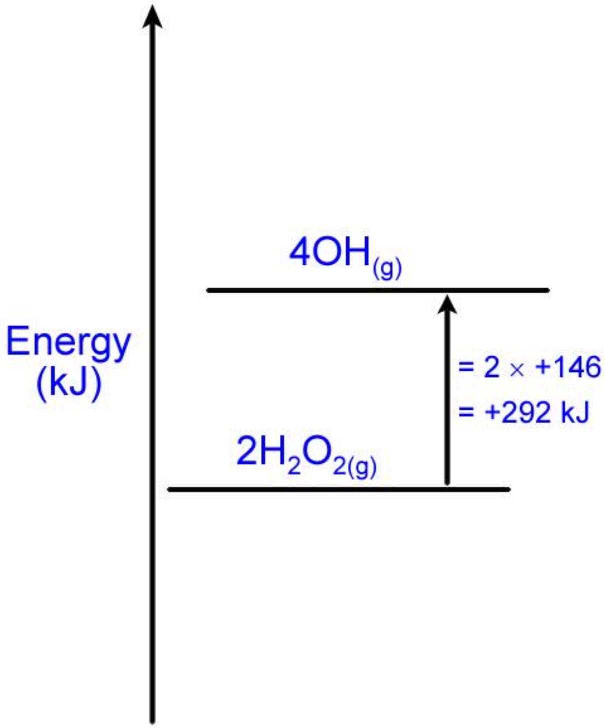

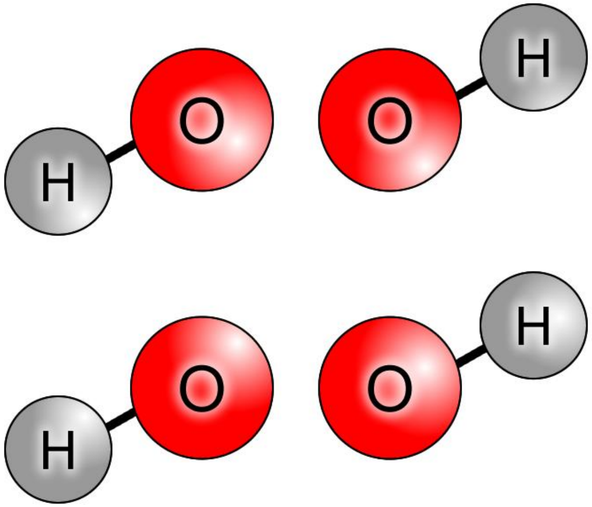

Bond breaking is *endothermic*: ΔH for this change is *positive*.

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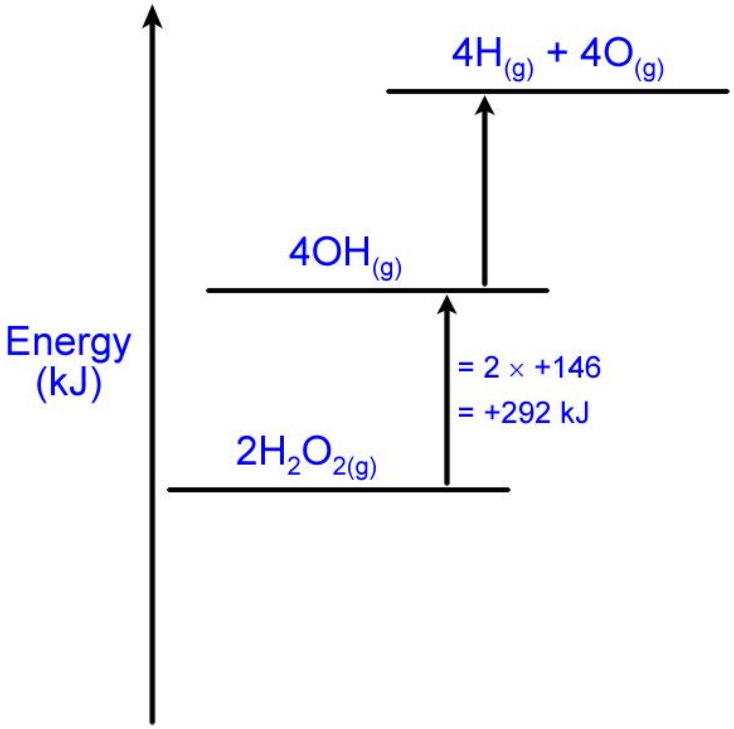
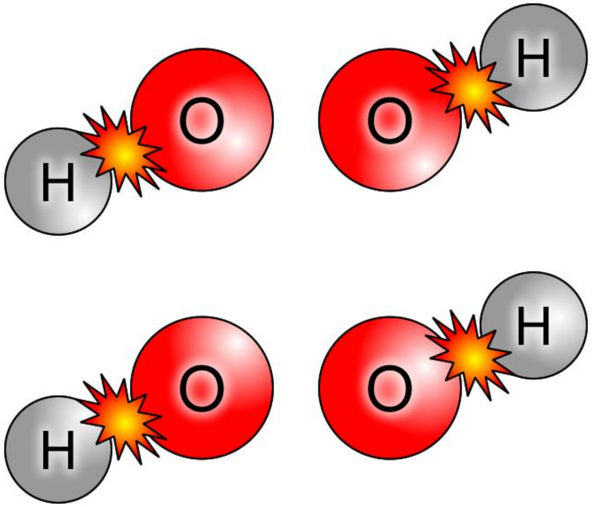




4 × O–H bonds are broken forming 4 × O and 4 × H atoms.

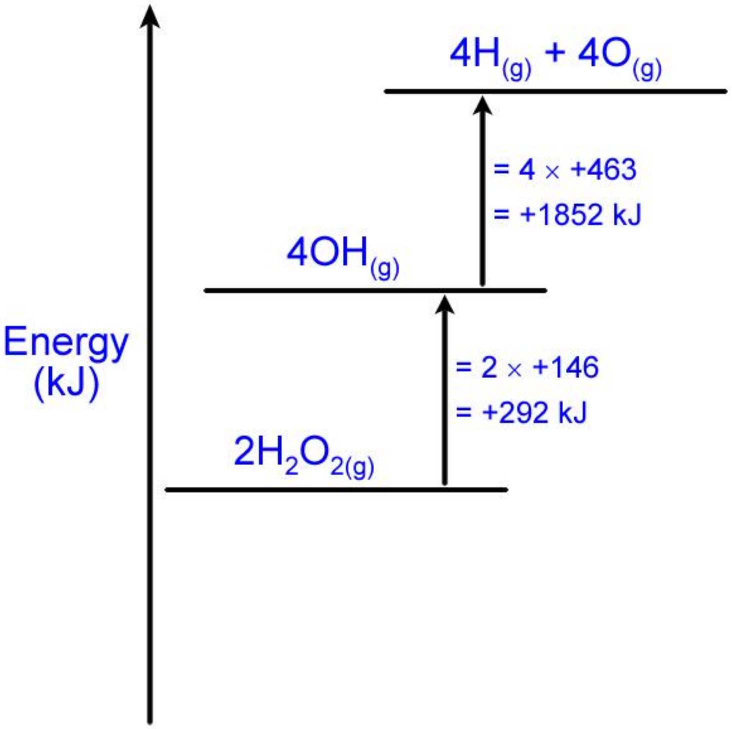
Enthalpy Change Calculations – Example #2:

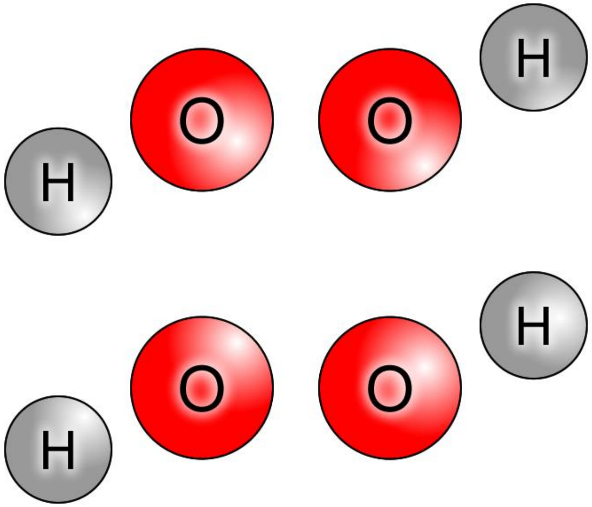
The Decomposition of Hydrogen Peroxide to form Oxygen and Water

$$2H_2O_2(g) \rightarrow O_2(g) + 2H_2O(g)$$

O–O = 146 kJ mol^{-1} O–H = 463 kJ mol^{-1} O=O = 496 kJ mol^{-1}

energy level diagram





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

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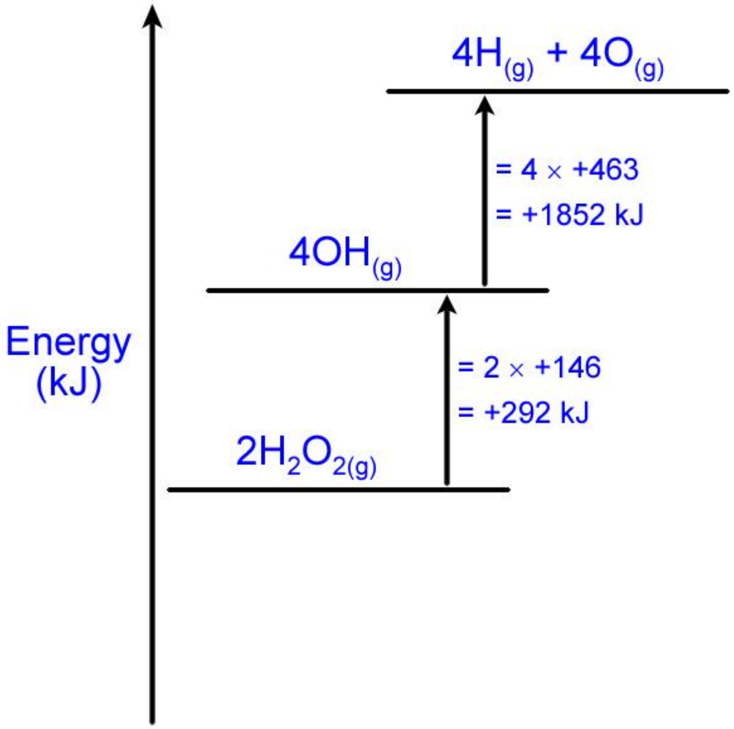

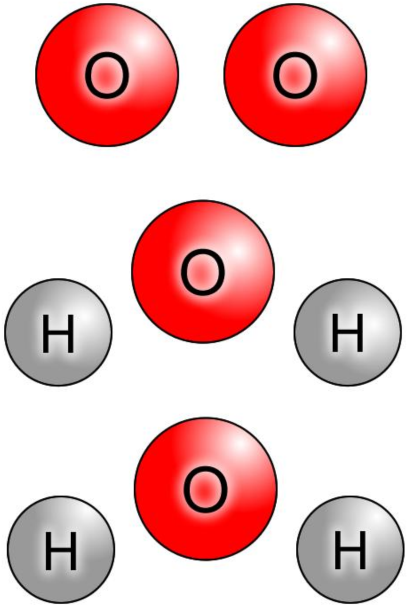

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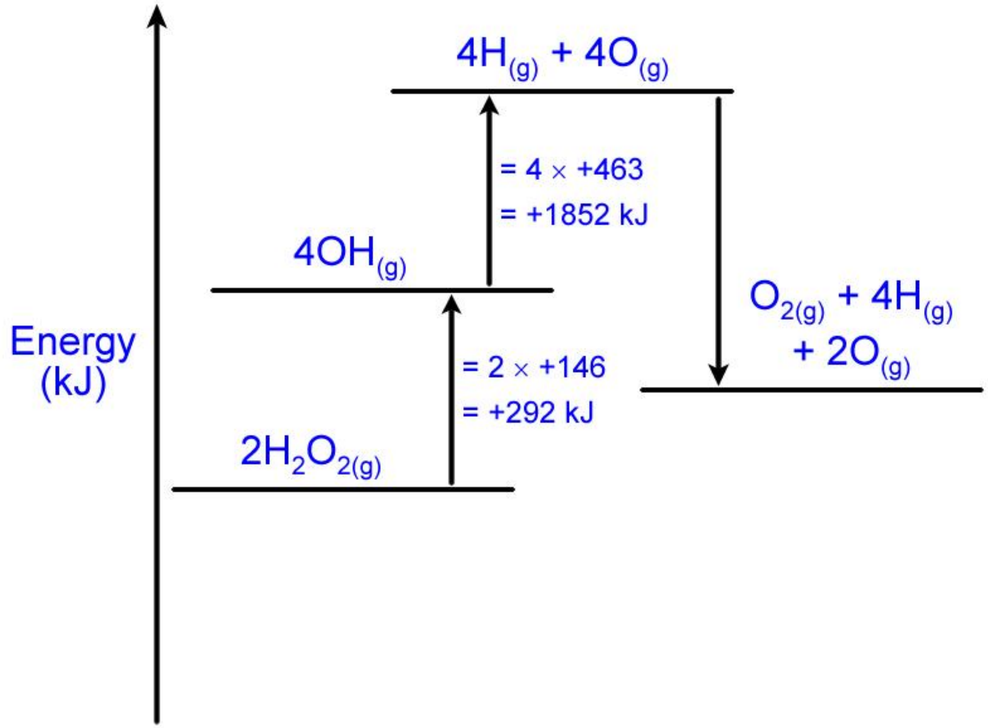

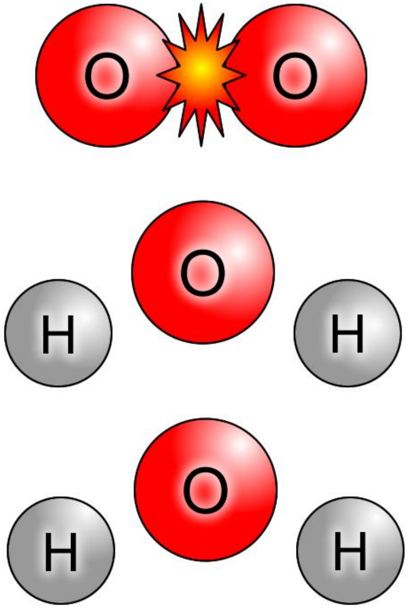

1 × O=O bond is formed creating 1 × O_2 molecule.

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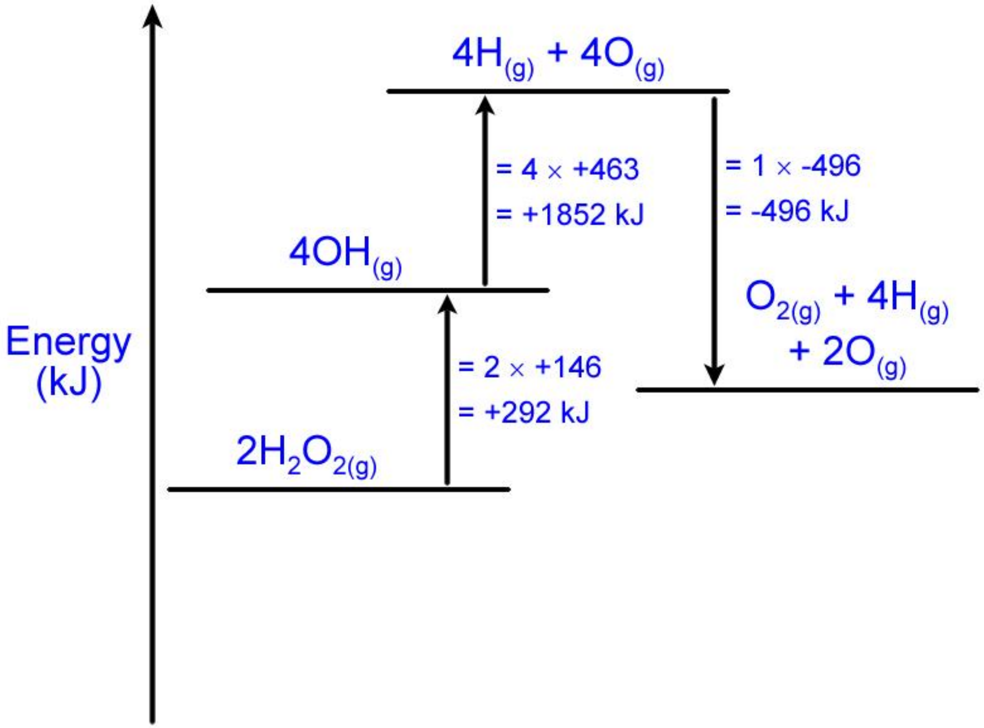

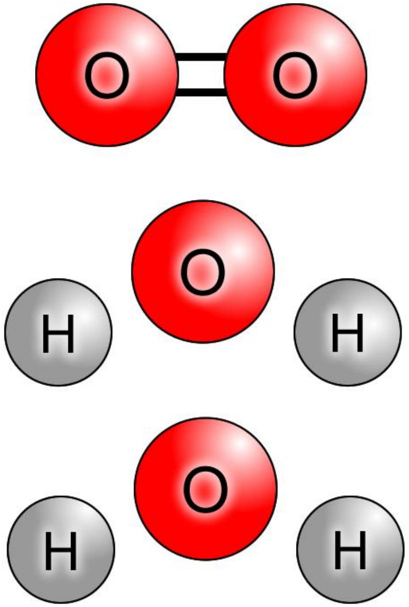

Bond formation is *exothermic*: ΔH for this change is *negative*.

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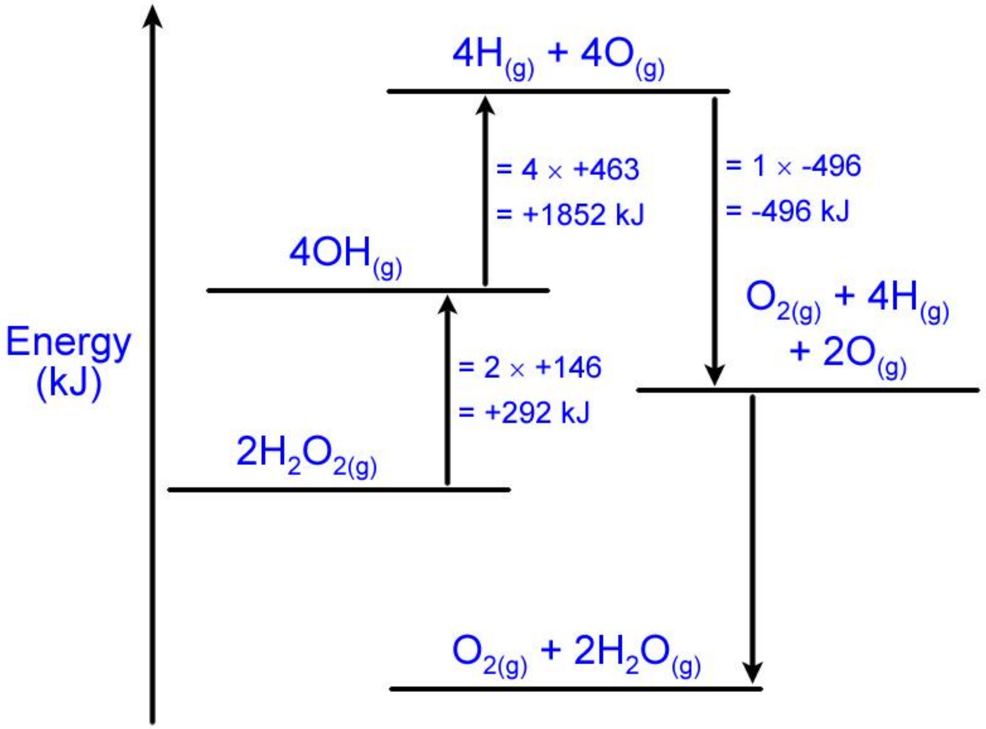

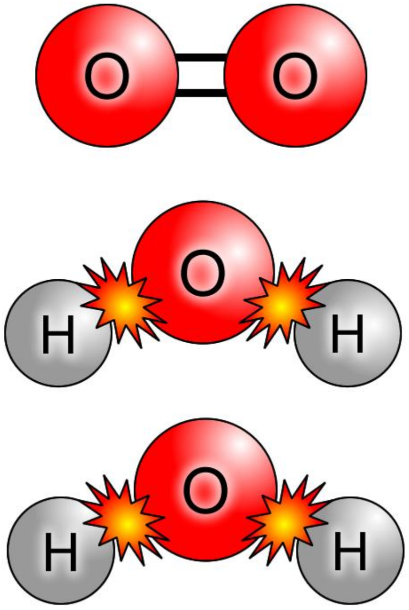

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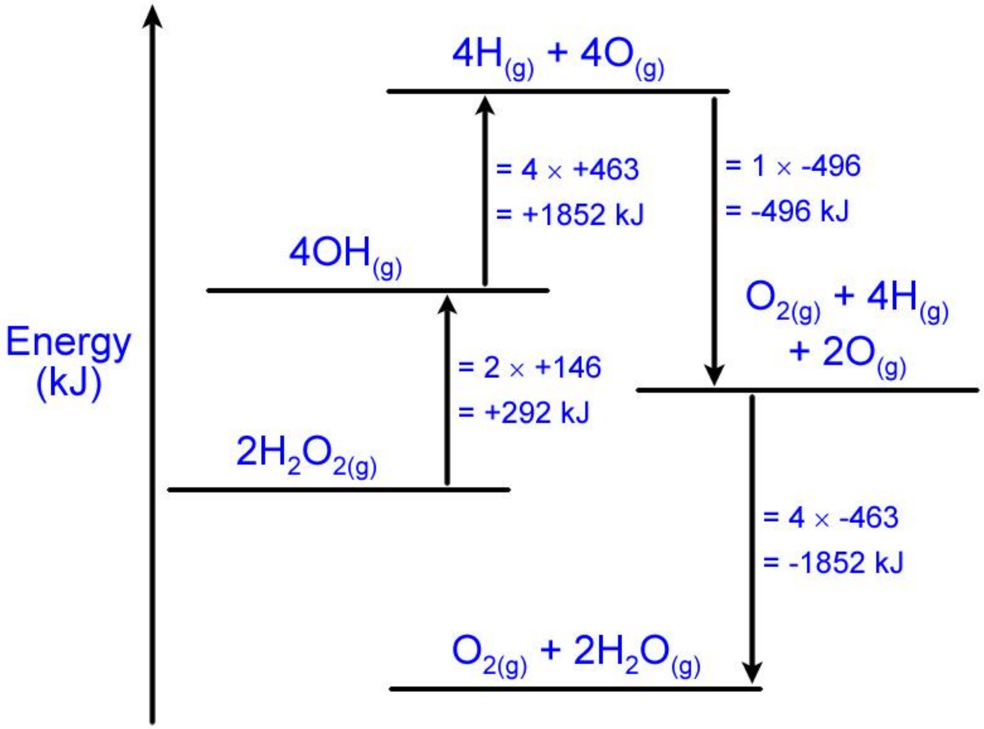

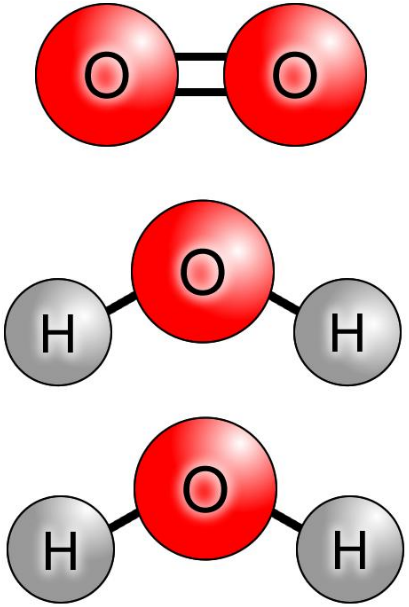

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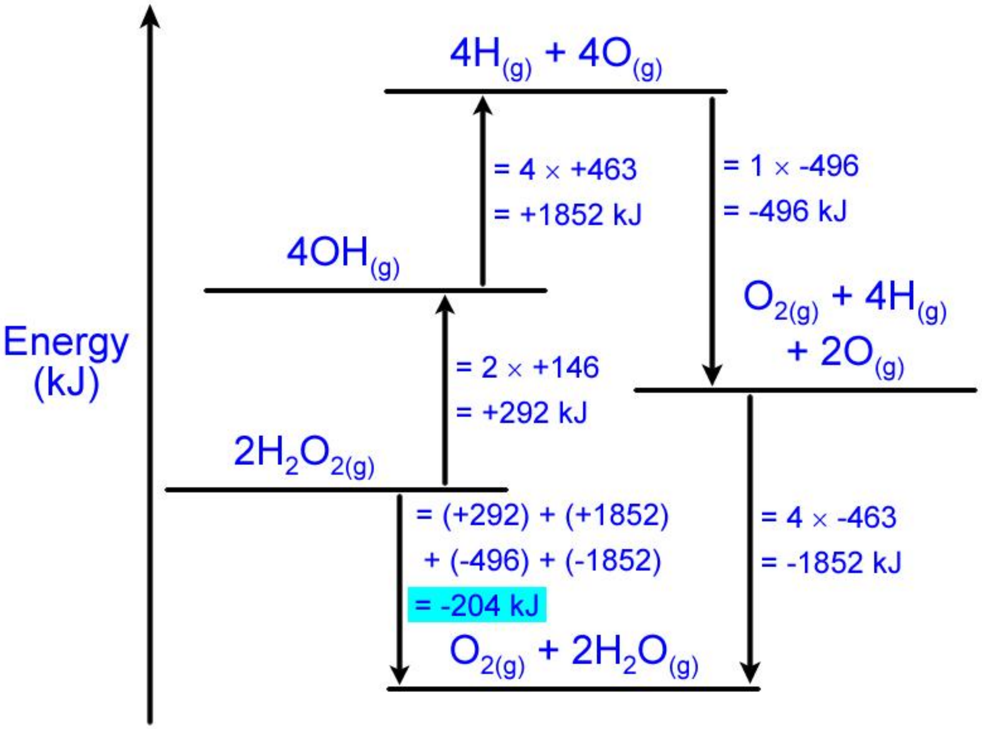

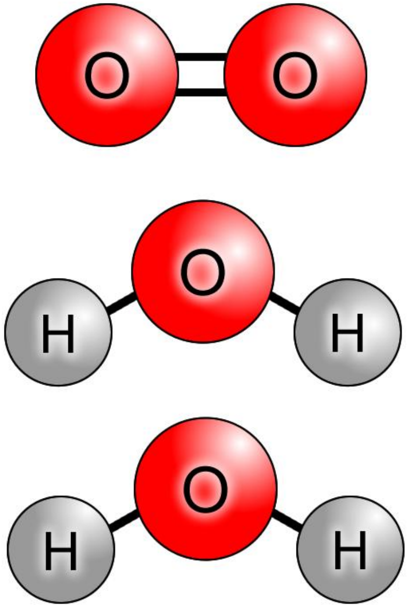

(+292) + (+1852) + (−496) + (−1852) = *−204 kJ mol^{-1}*

Enthalpy Change Calculations – Example #2:
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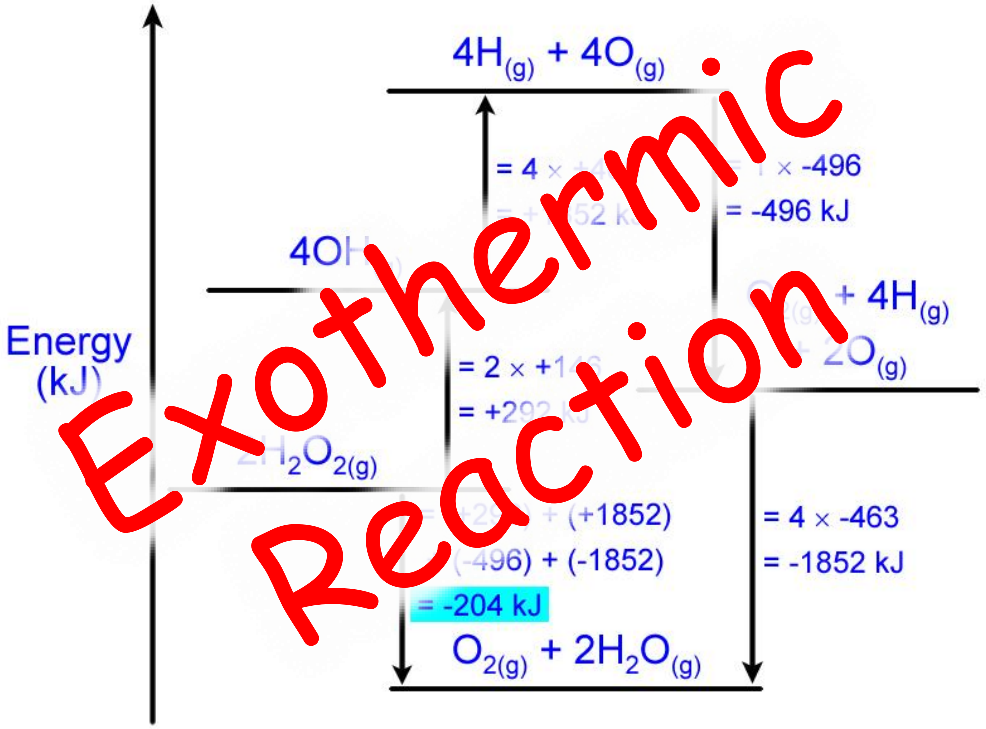

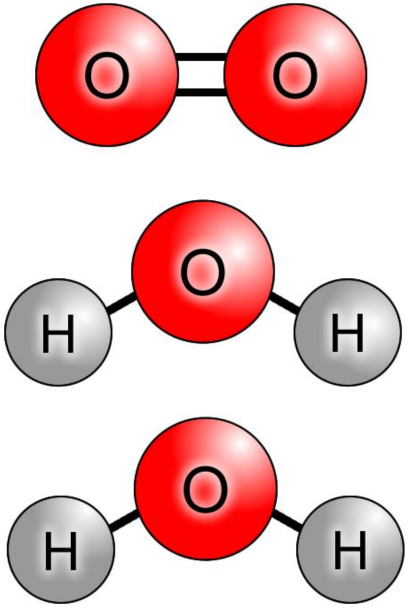

$(+292) + (+1852) + (-496) + (-1852) = \boxed{-204\ \text{kJ mol}^{-1}}$

Enthalpy Change Calculations – Example #3:

The Combustion of Methane forming Carbon Dioxide and Water

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$

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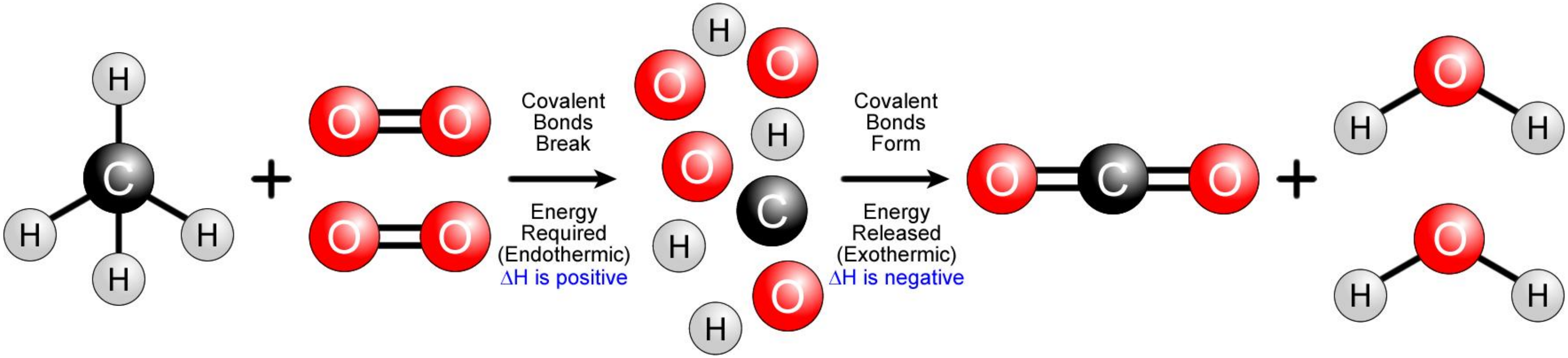

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$$CH_4(g) \ + \ 2O_2(g) \ \rightarrow \ CO_2(g) \ + \ 2H_2O(g)$$

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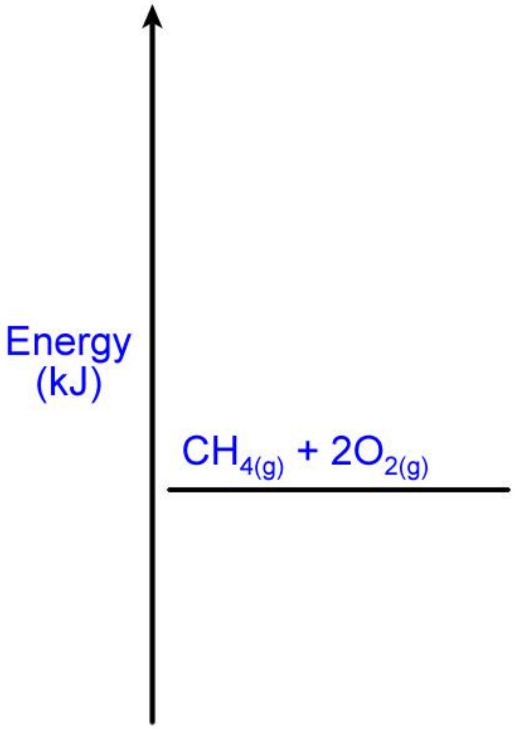

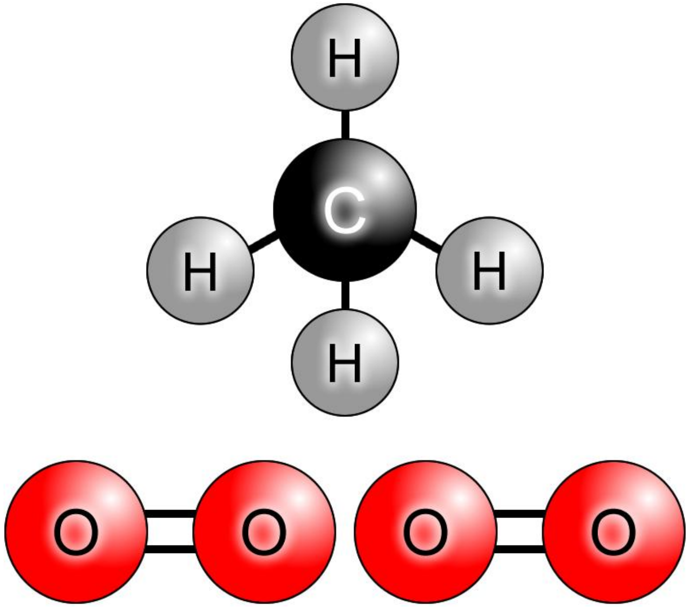

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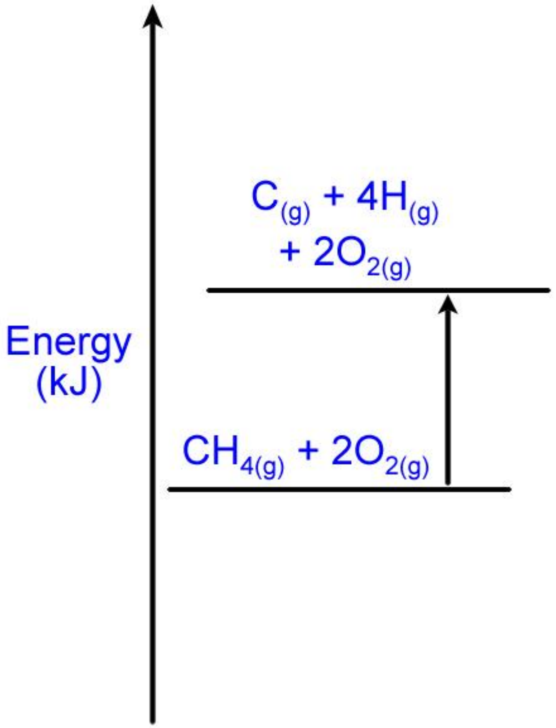

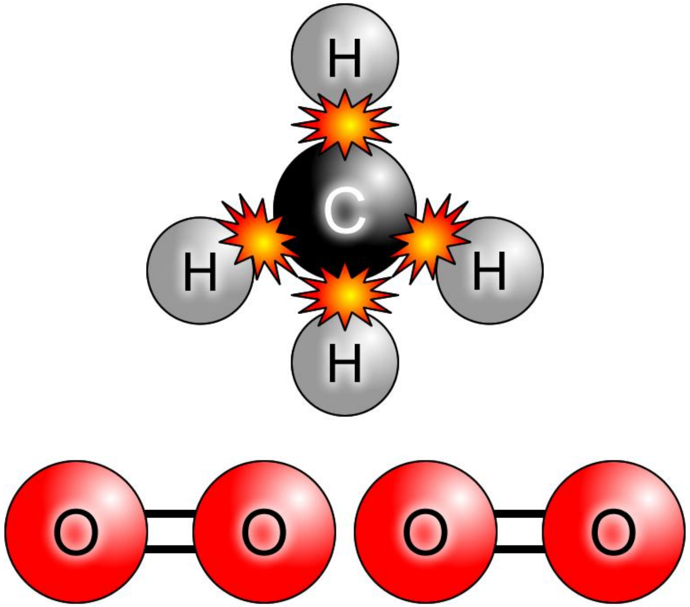

4 × C–H bonds are broken forming 1 × C atom and 4 × H atoms.

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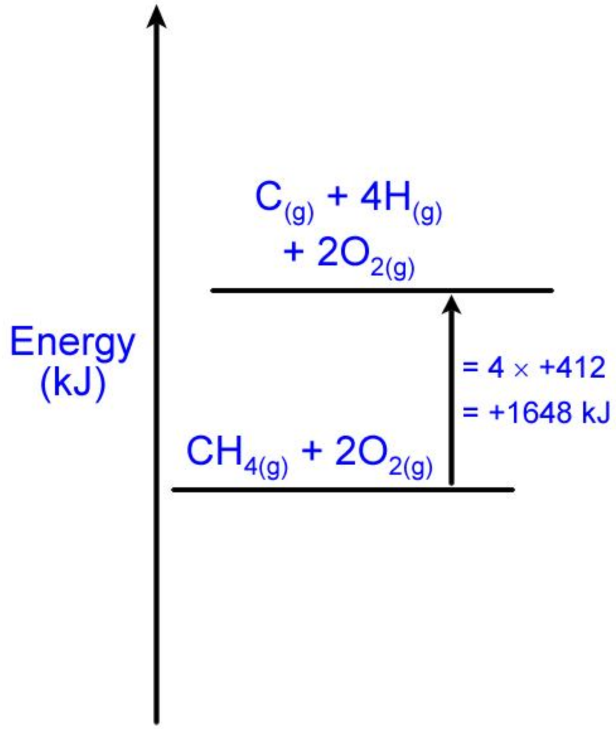

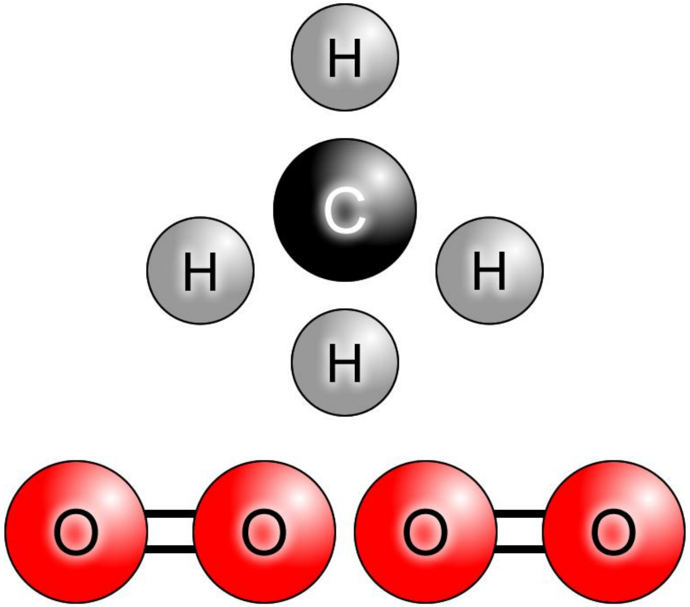

Bond breaking is *endothermic*: ΔH for this change is *positive*.

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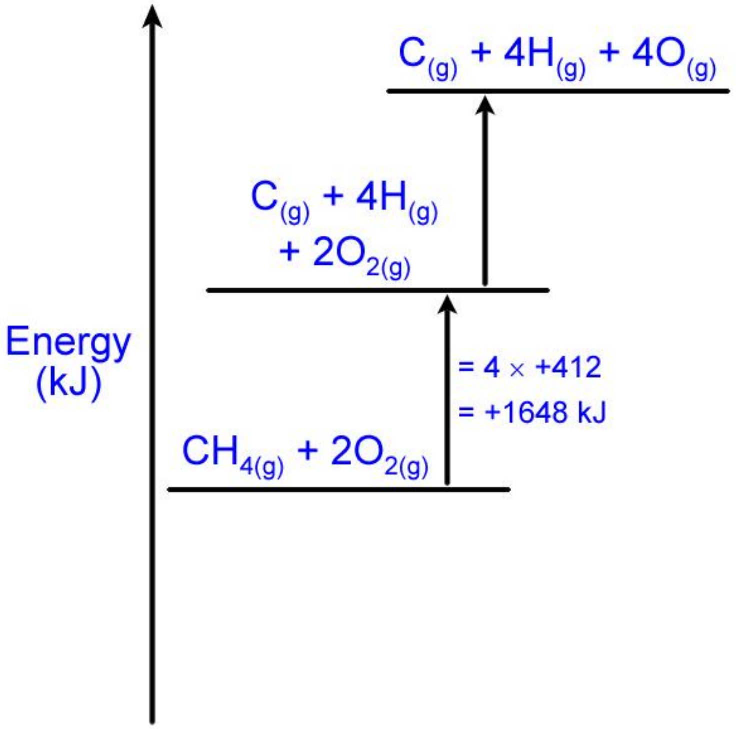

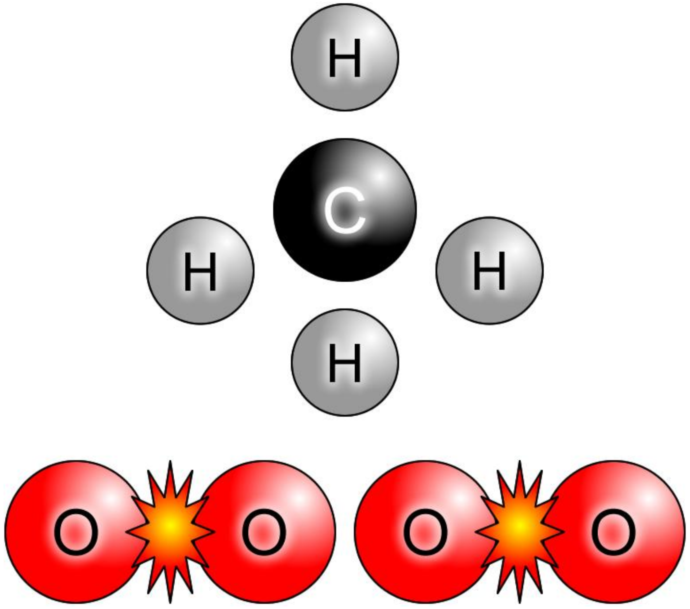

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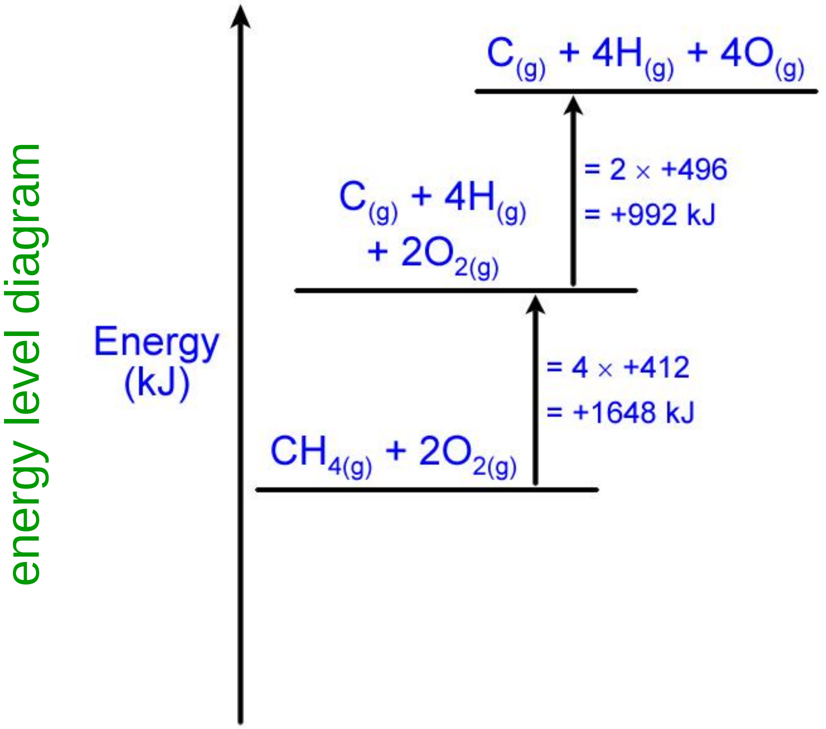

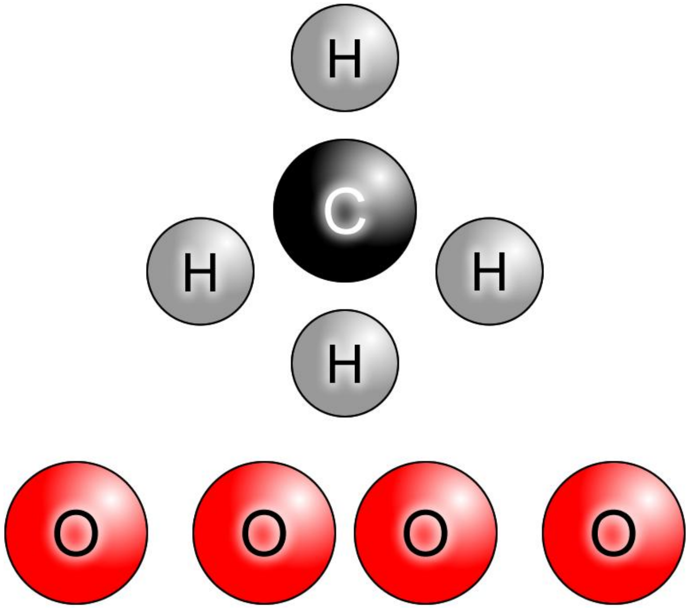

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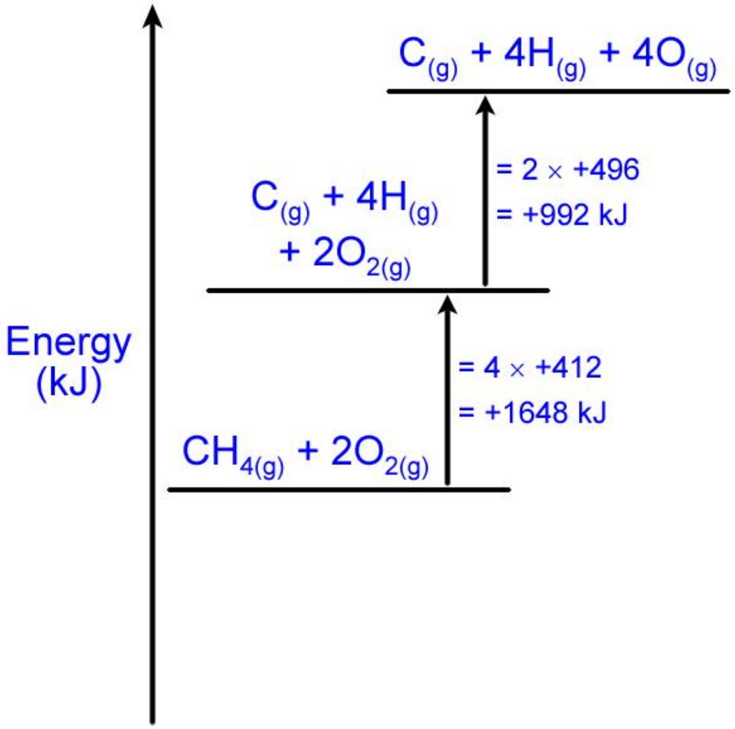

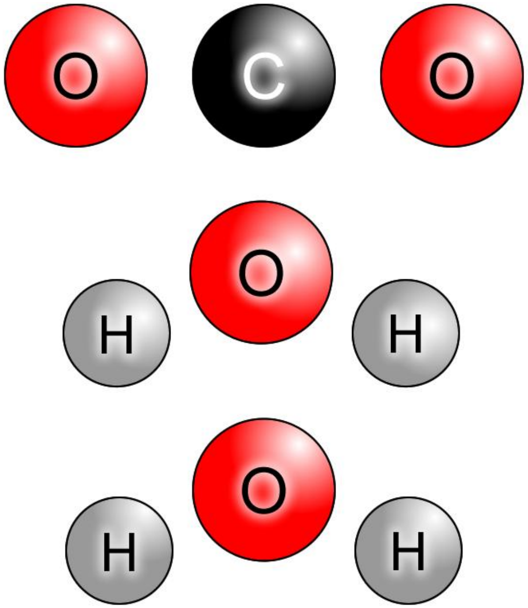

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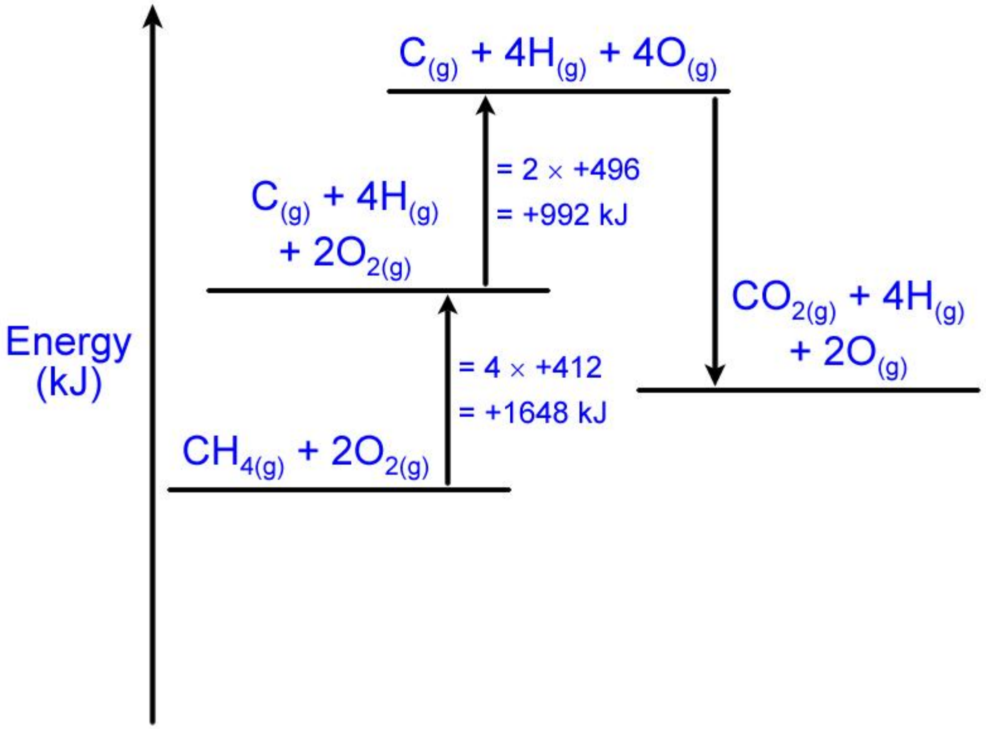

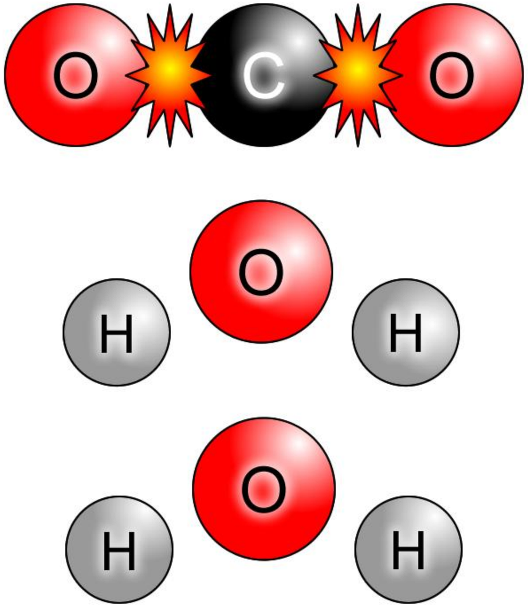

2 × C=O bonds are formed creating 1 × CO_2 molecule.

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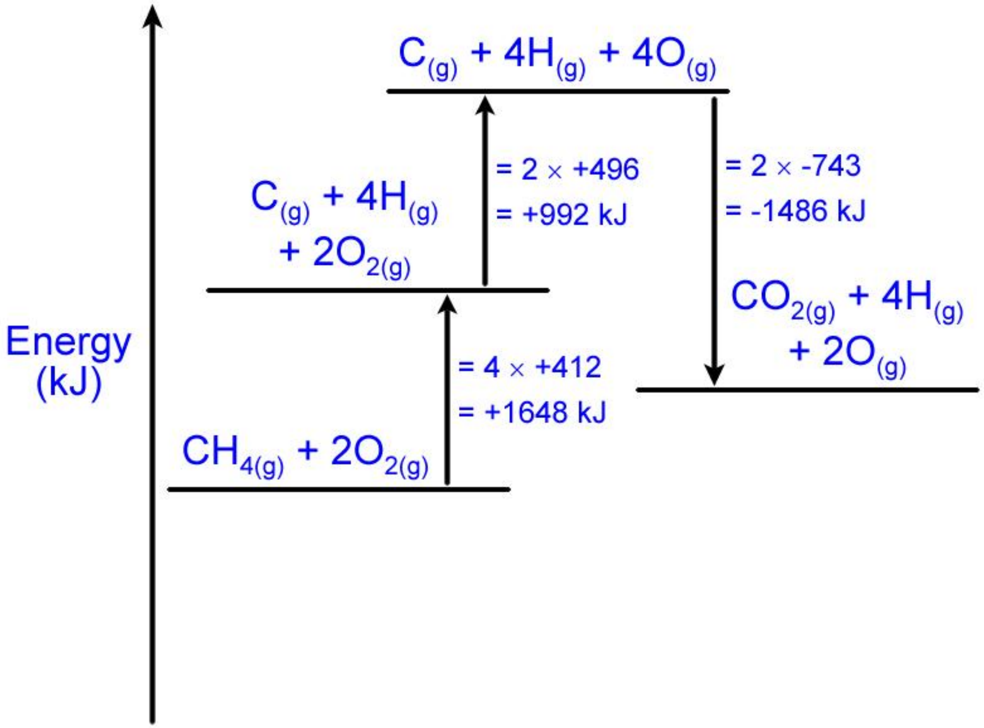

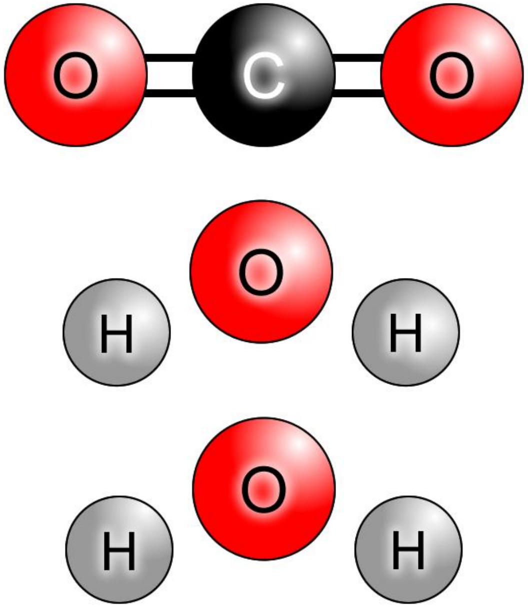

Bond formation is *exothermic*: ΔH for this change is *negative*.

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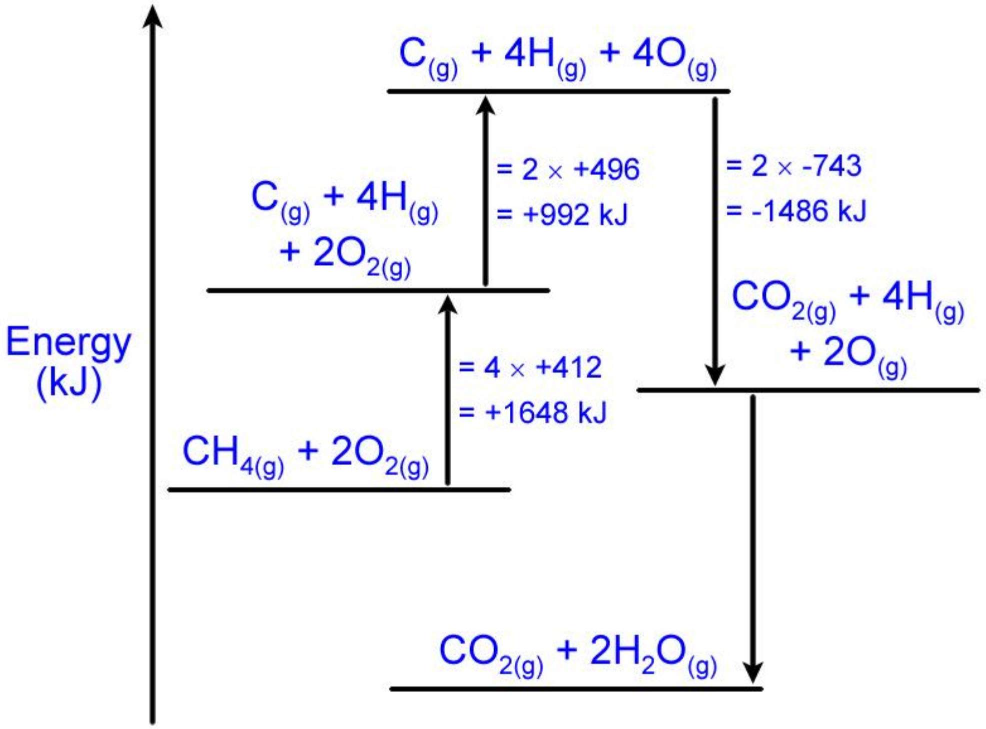

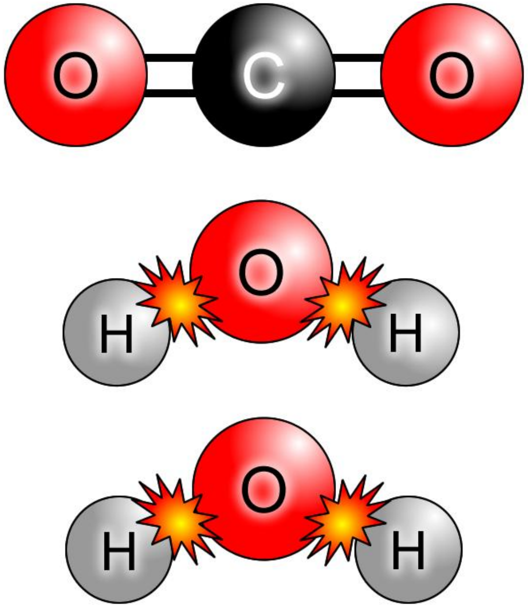

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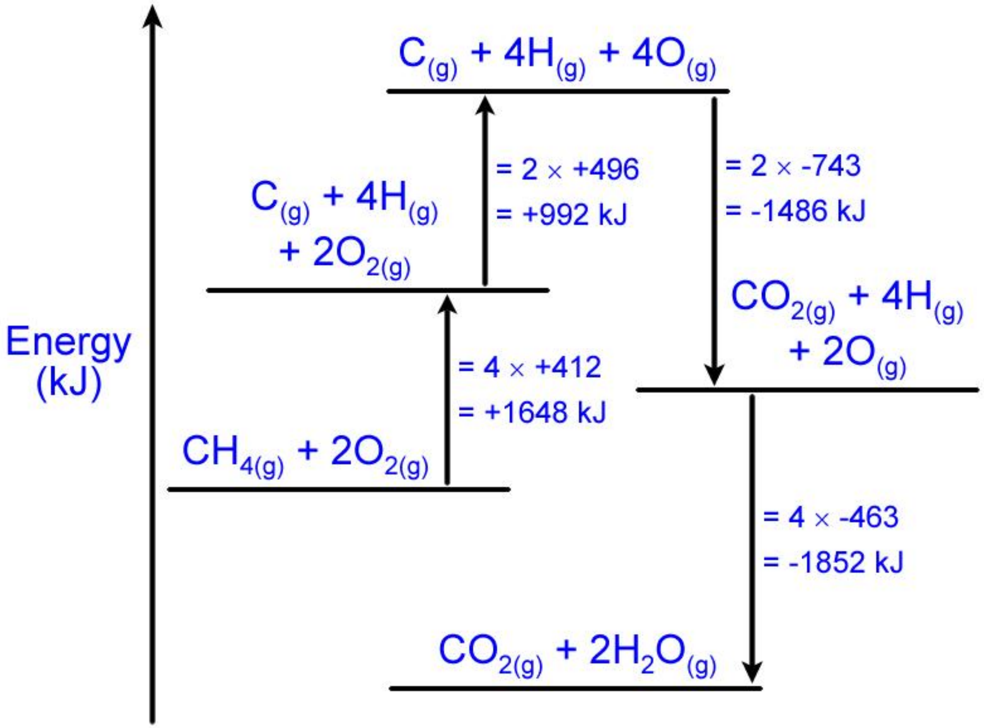

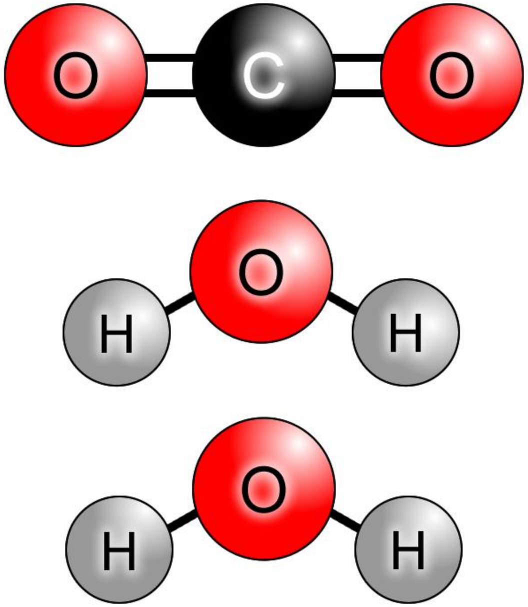

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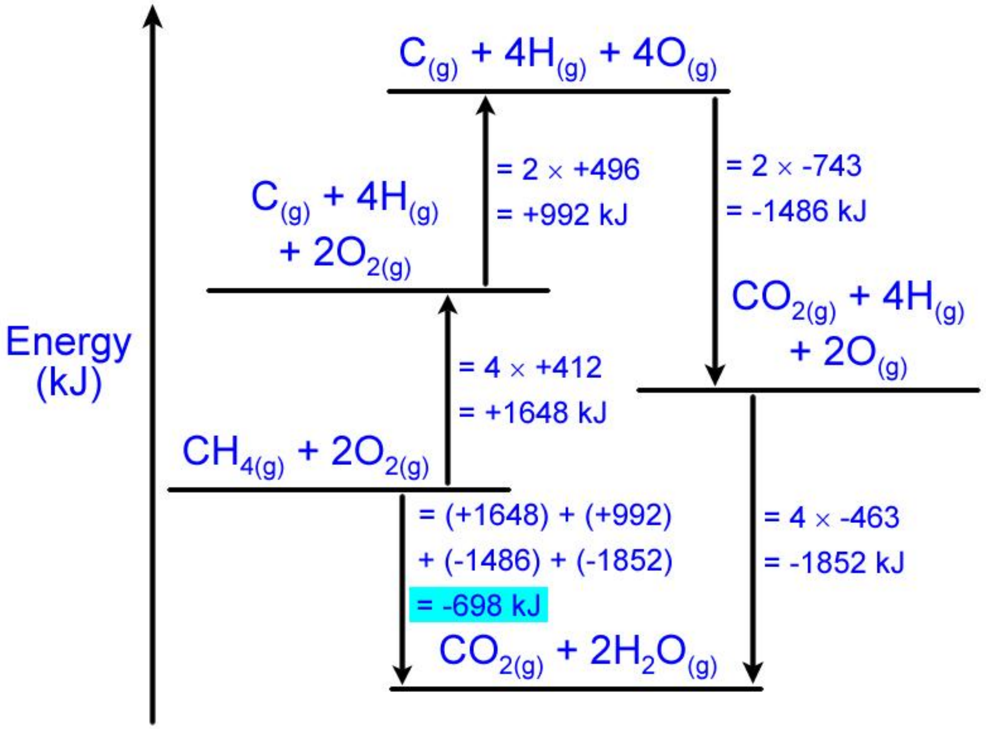

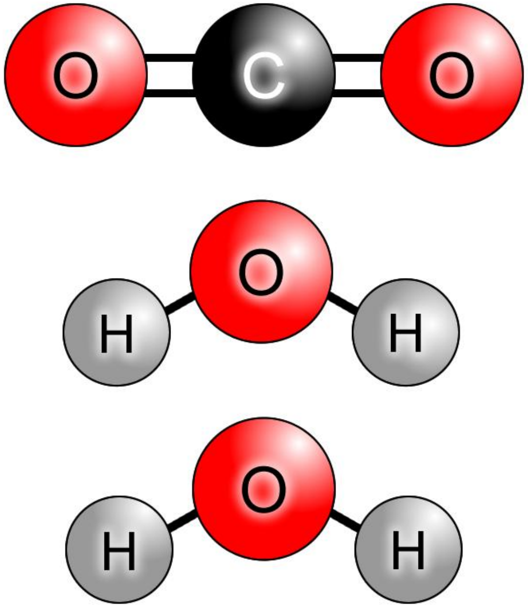

$$(+1648) + (+992) + (-1486) + (-1852) = -698\ \text{kJ mol}^{-1}$$

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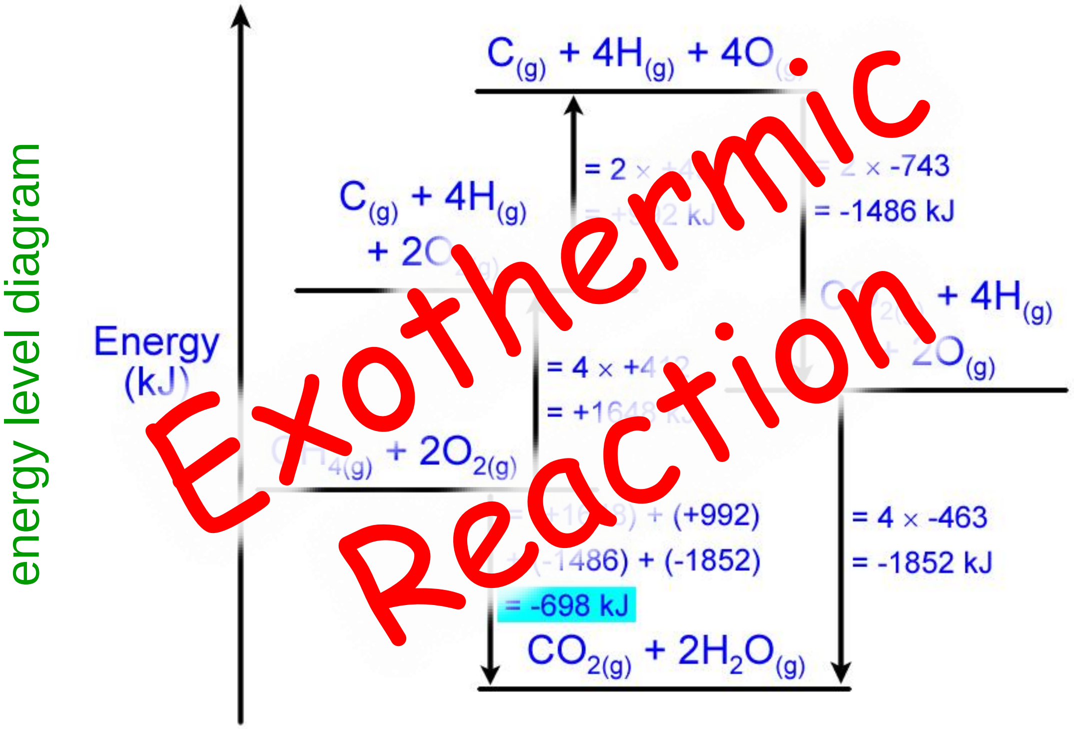
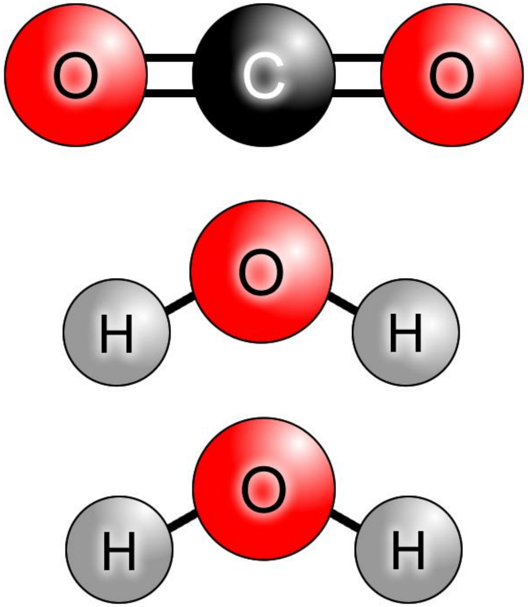




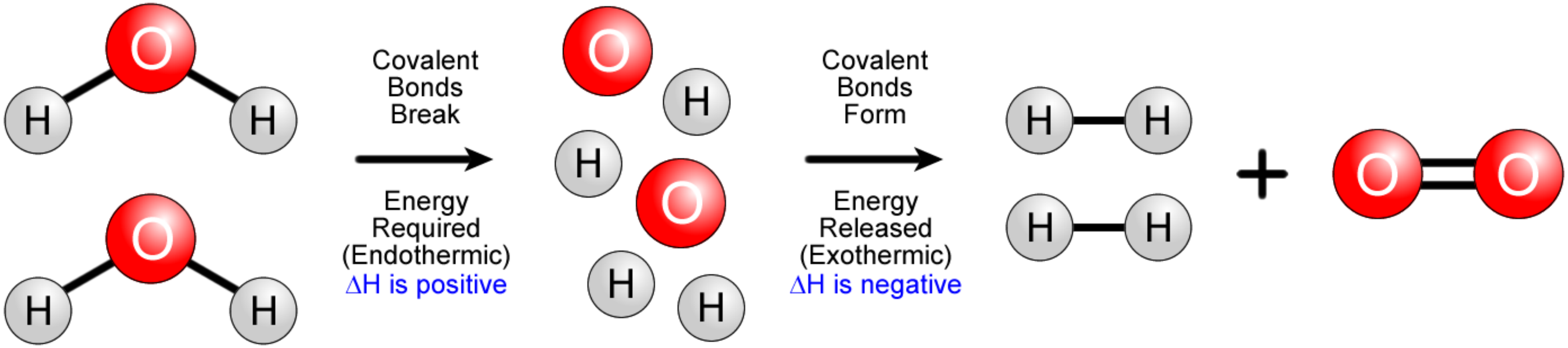
(+1648) + (+992) + (−1486) + (−1852) = *−698 kJ mol^{-1}*

Enthalpy 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

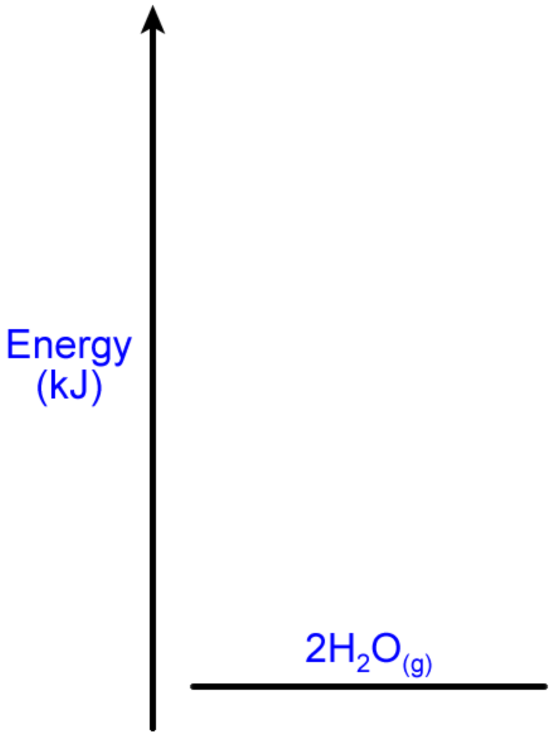
Enthalpy Change Calculations – Example #4:

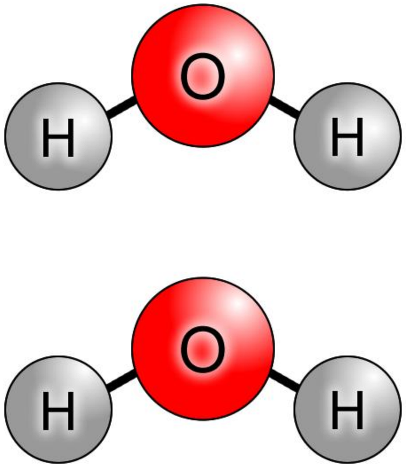
The Electrolysis of Water to form Hydrogen and Oxygen

$$2H_2O(g) \rightarrow 2H_2(g) + O_2(g)$$

O–H = 463 kJ mol^{-1} O=O = 496 kJ mol^{-1} H–H = 436 kJ mol^{-1}

energy level diagram





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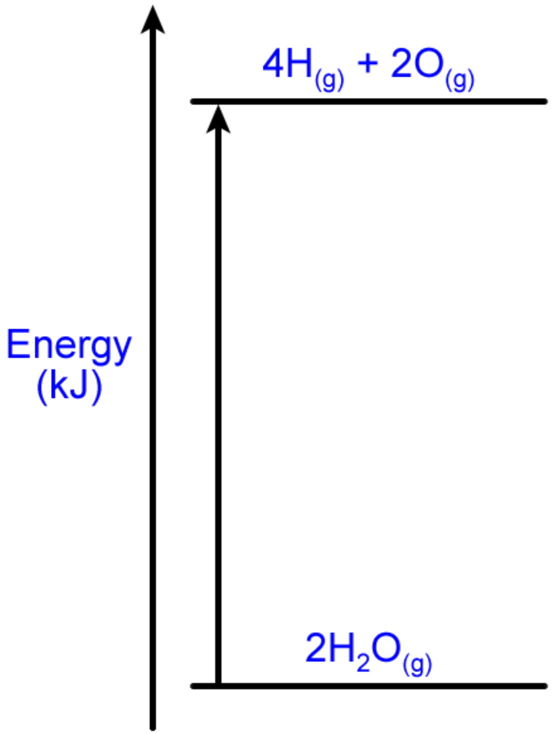

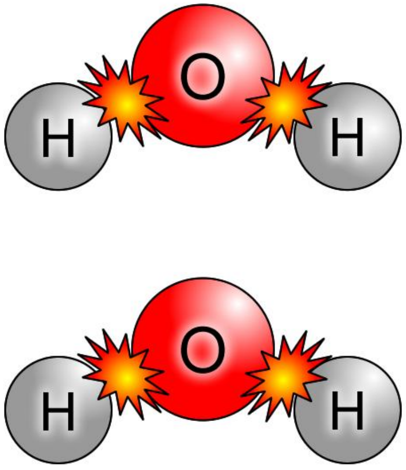

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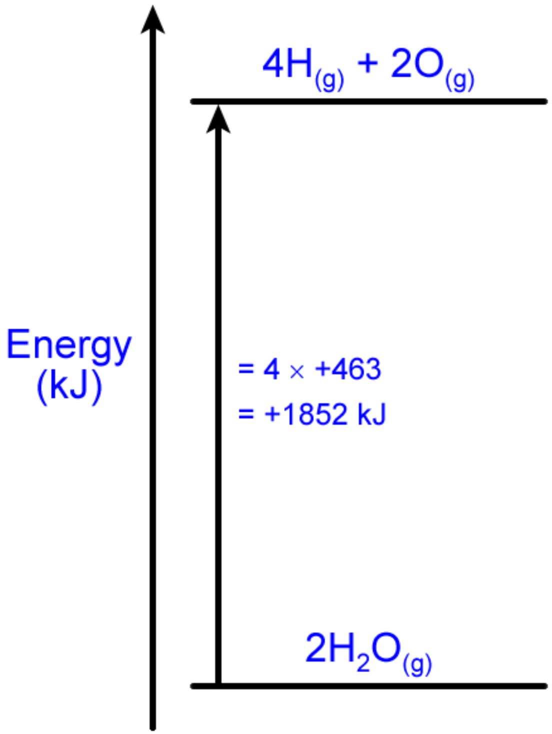
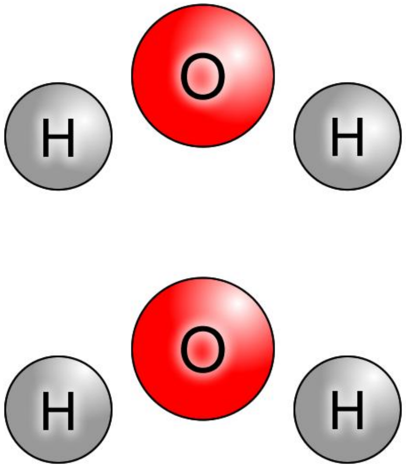
O–H = 463 kJ mol^{-1} O=O = 496 kJ mol^{-1} H–H = 436 kJ mol^{-1}

energy level diagram





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

Enthalpy 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^{-1} O=O = 496 kJ mol^{-1} H–H = 436 kJ mol^{-1}

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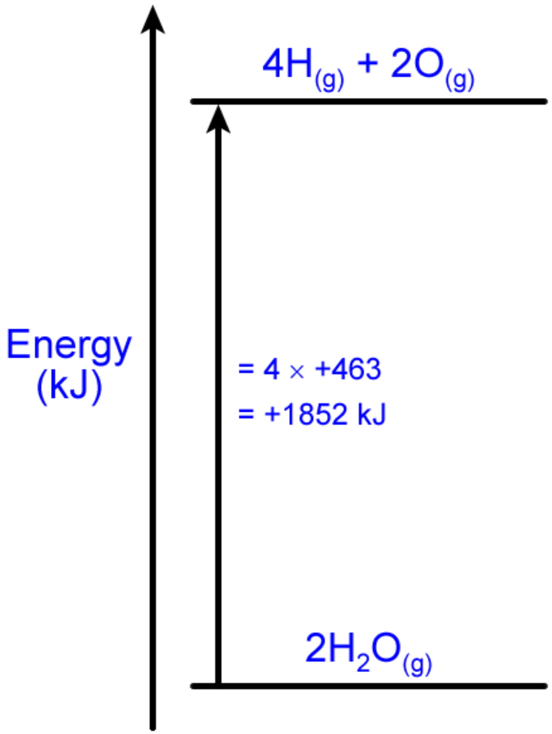

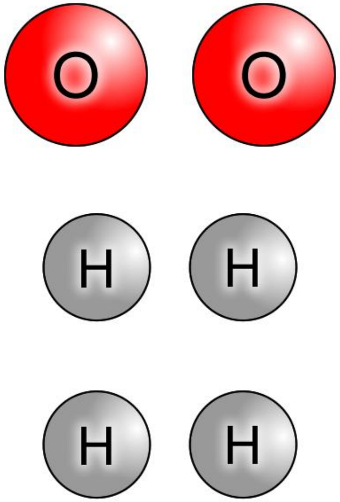

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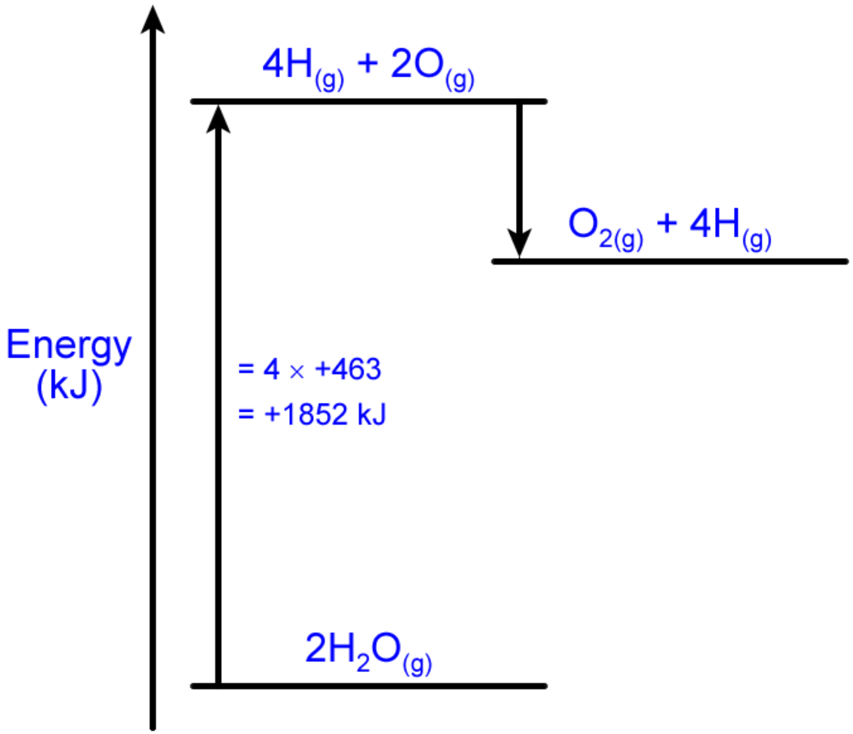
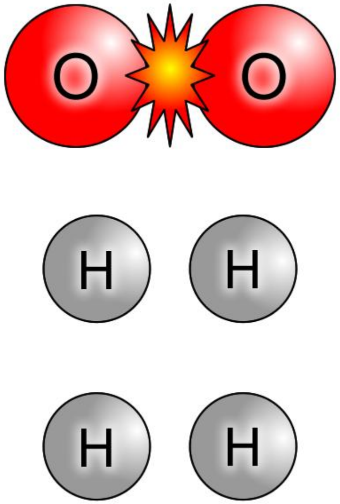


energy level diagram



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

Enthalpy Change Calculations – Example #4:

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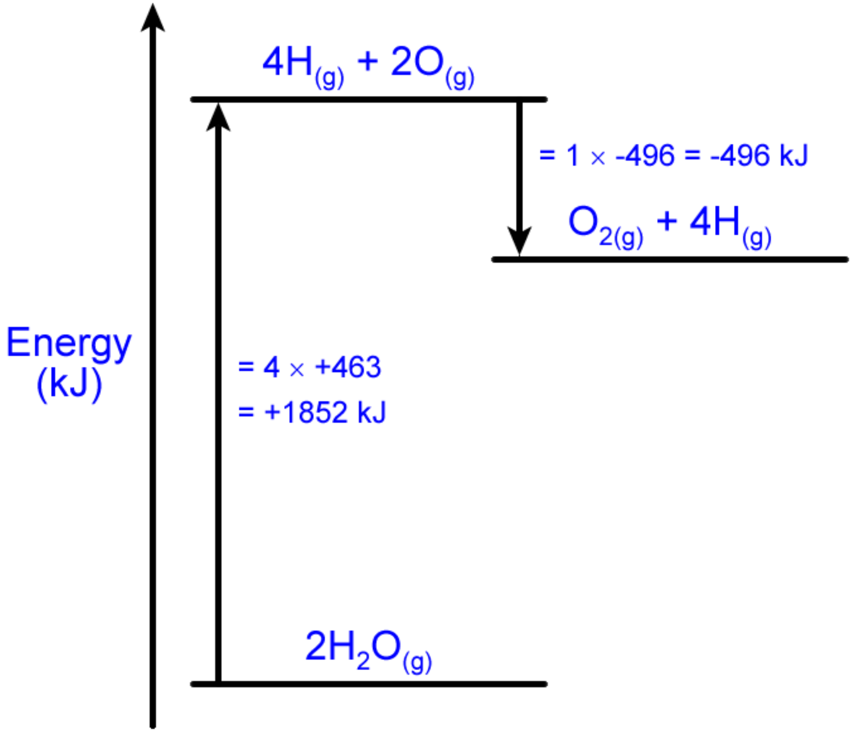


Energy (kJ)

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

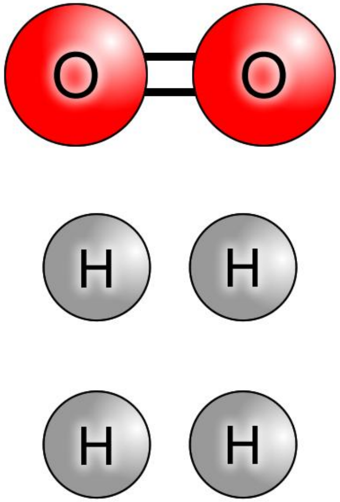
= 1 × -496 = -496 kJ

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

= 4 × +463

= +1852 kJ

$2H_2O_{(g)}$



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

Enthalpy 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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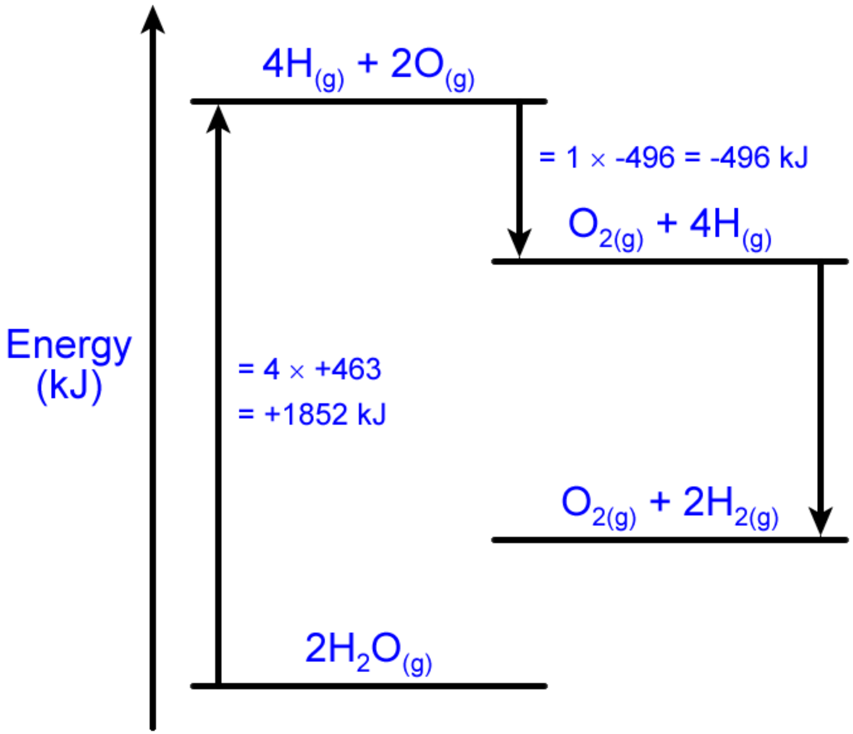

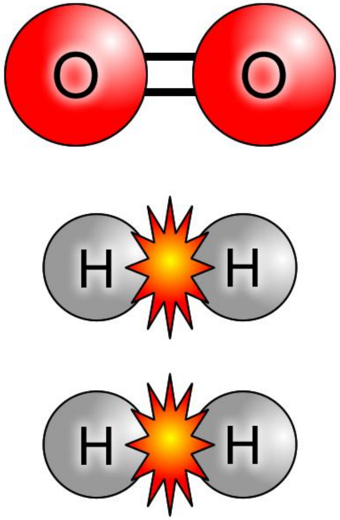

2 × H–H bonds are formed creating 2 × H_2 molecules.

Enthalpy 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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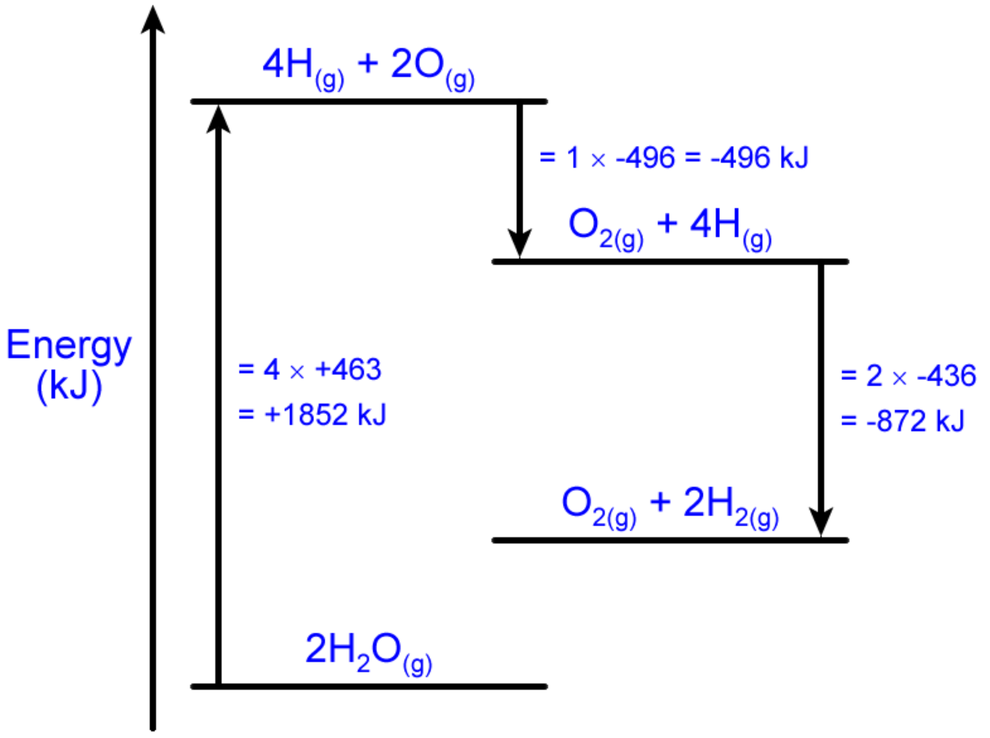

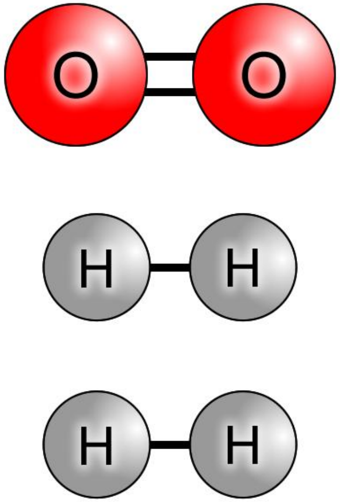

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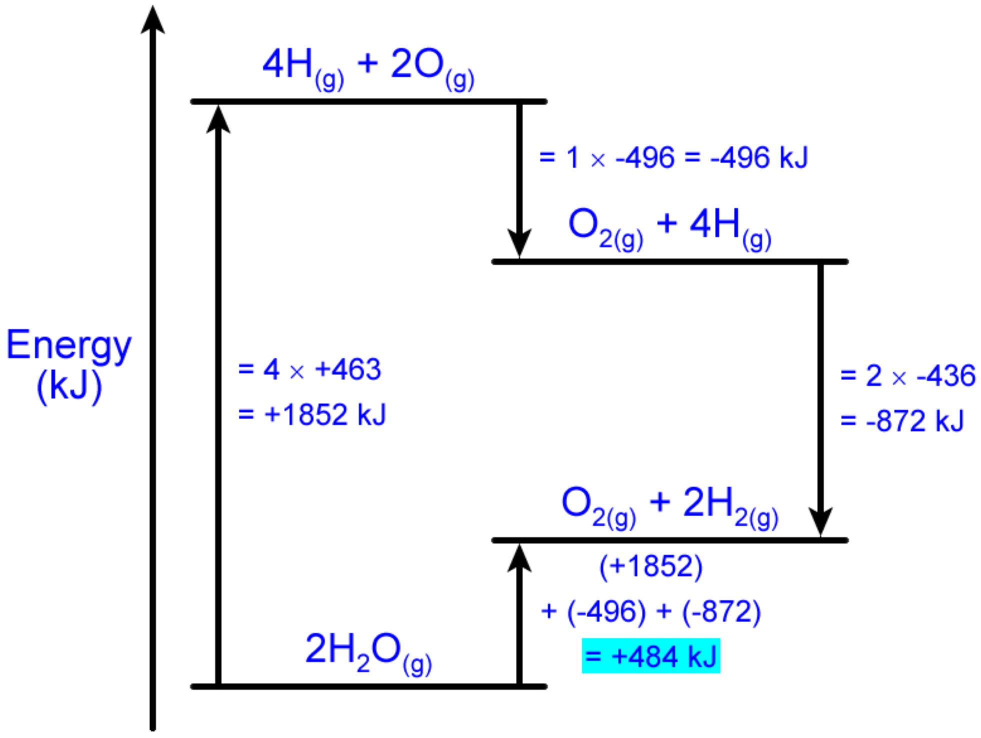


Energy (kJ)

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

= 1 × -496 = -496 kJ

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

= 4 × +463

= +1852 kJ

= 2 × -436

= -872 kJ

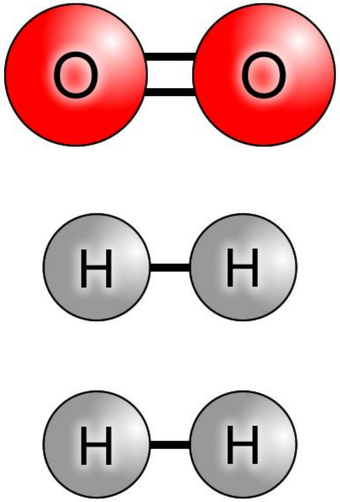
$O_{2(g)} + 2H_{2(g)}$

(+1852)

+ (-496) + (-872)

= +484 kJ

$2H_2O_{(g)}$



$$(+1852) + (-496) + (-872) = +484 \text{ kJ mol}^{-1}$$

Enthalpy Change Calculations – Example #4:

The Electrolysis of Water to form Hydrogen and Oxygen

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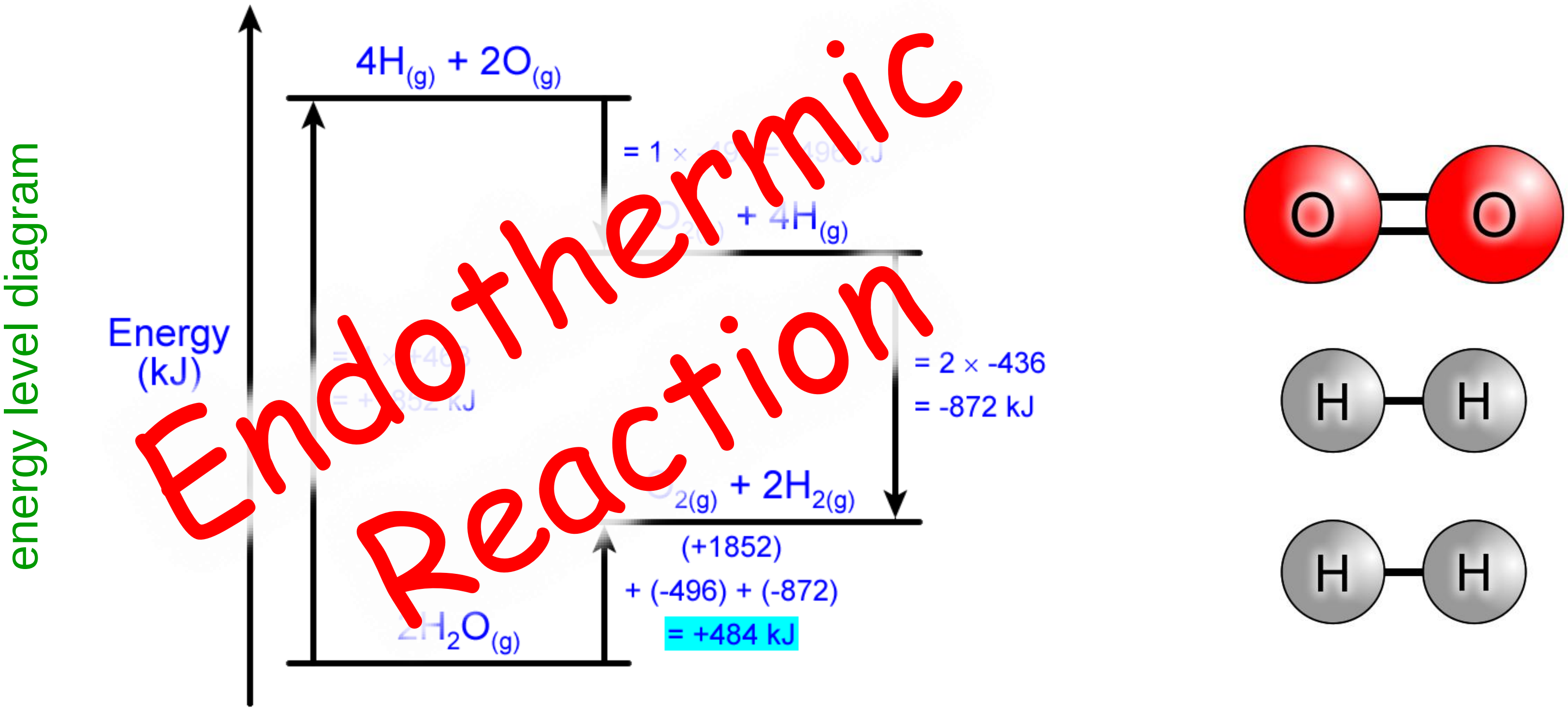
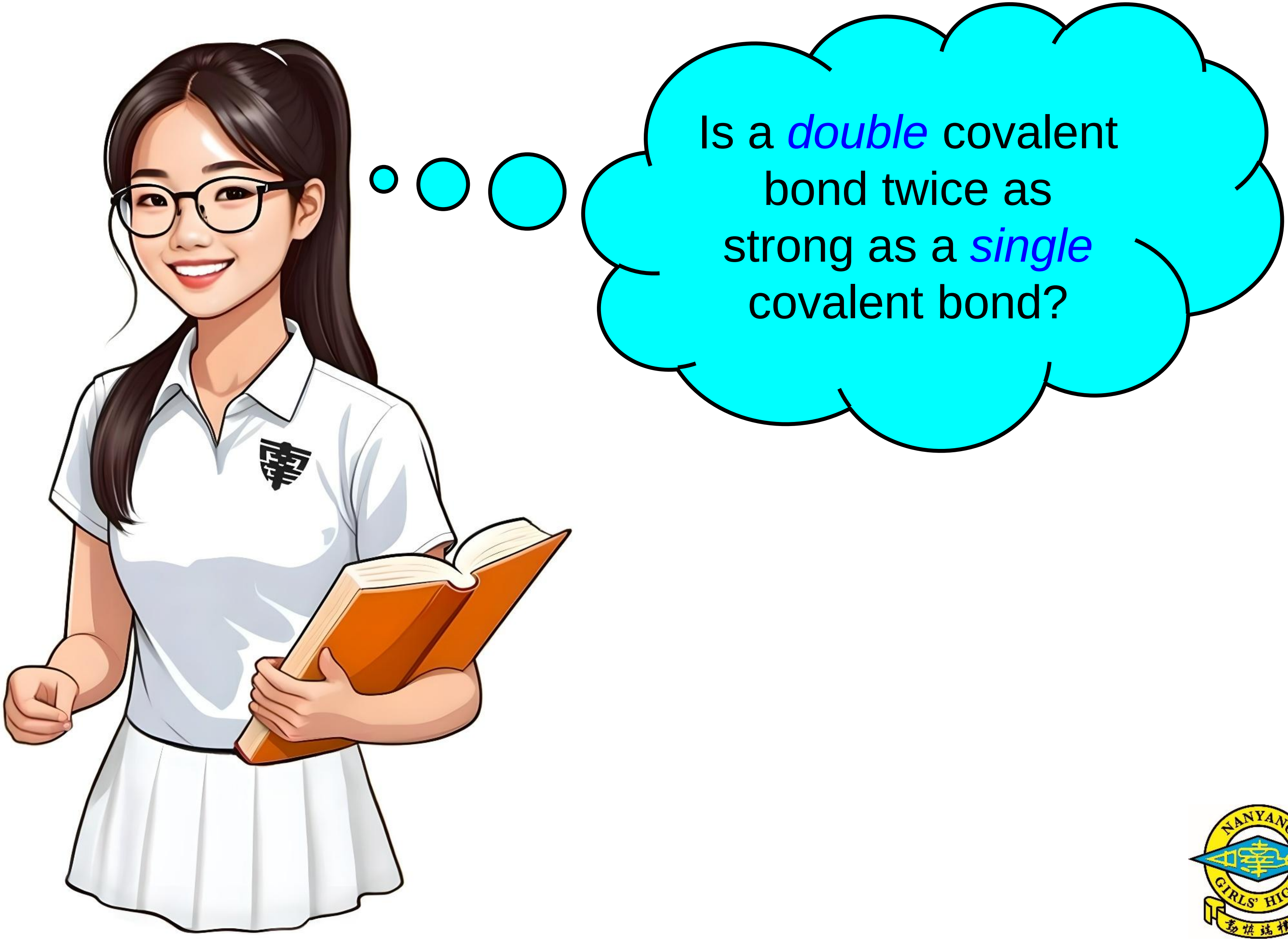


$(+1852) + (-496) + (-872) = \boxed{+484\ kJ\ mol^{-1}}$

Energy from Chemicals – Enthalpy Changes



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

Energy from Chemicals – Enthalpy Changes

- Study the bond energies given below:

C–C = 348 kJ C=C = 612 kJ

O–O = 146 kJ O=O = 496 kJ

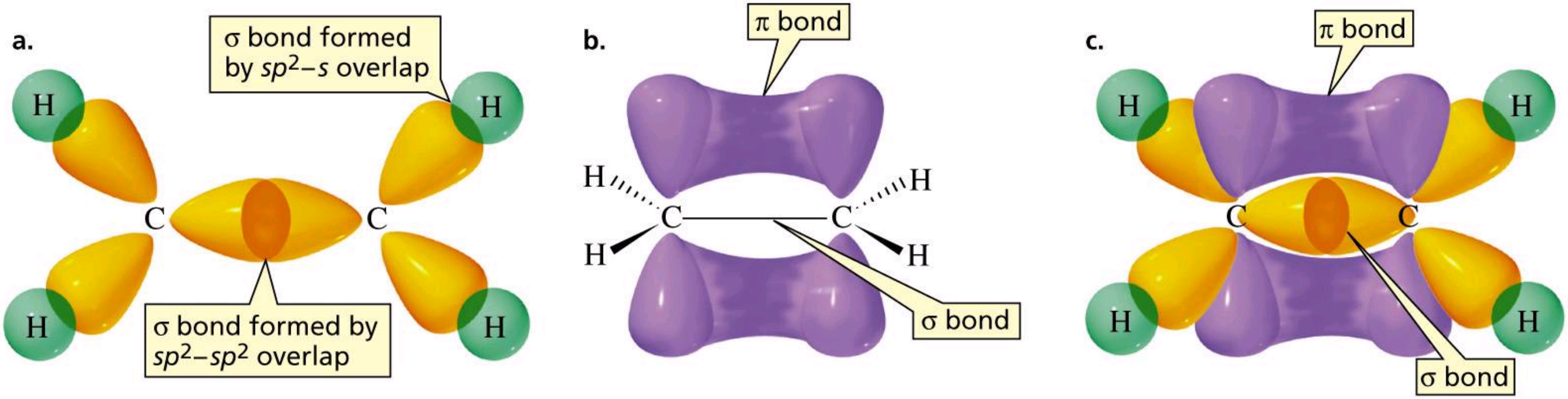
- Is a double covalent bond twice as strong as a single covalent bond?

- 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 – Enthalpy Changes

- 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 – Enthalpy of Solution

What affects the enthalpy change when a chemical *dissolves in water*?

Energy from Chemicals – Enthalpy of Solution

- The enthalpy change of solution is the enthalpy change associated with dissolving one mole of a chemical in a solvent (usually water) at room temperature and pressure.

- The enthalpy change of solution can be *exothermic*, *e.g.* dissolving *sodium hydroxide* in distilled water:

$$NaOH(s) \rightarrow Na^{+}(aq) + OH^{-}(aq) \qquad \Delta H = -45.0 \text{ kJ mol}^{-1}$$

- The enthalpy change of solution can be *endothermic*, *e.g.* dissolving *ammonium nitrate* in distilled water:

$$NH_4NO_3(s) \rightarrow NH_4^{+}(aq) + NO_3^{-}(aq) \qquad \Delta H = +25.0 \text{ kJ mol}^{-1}$$

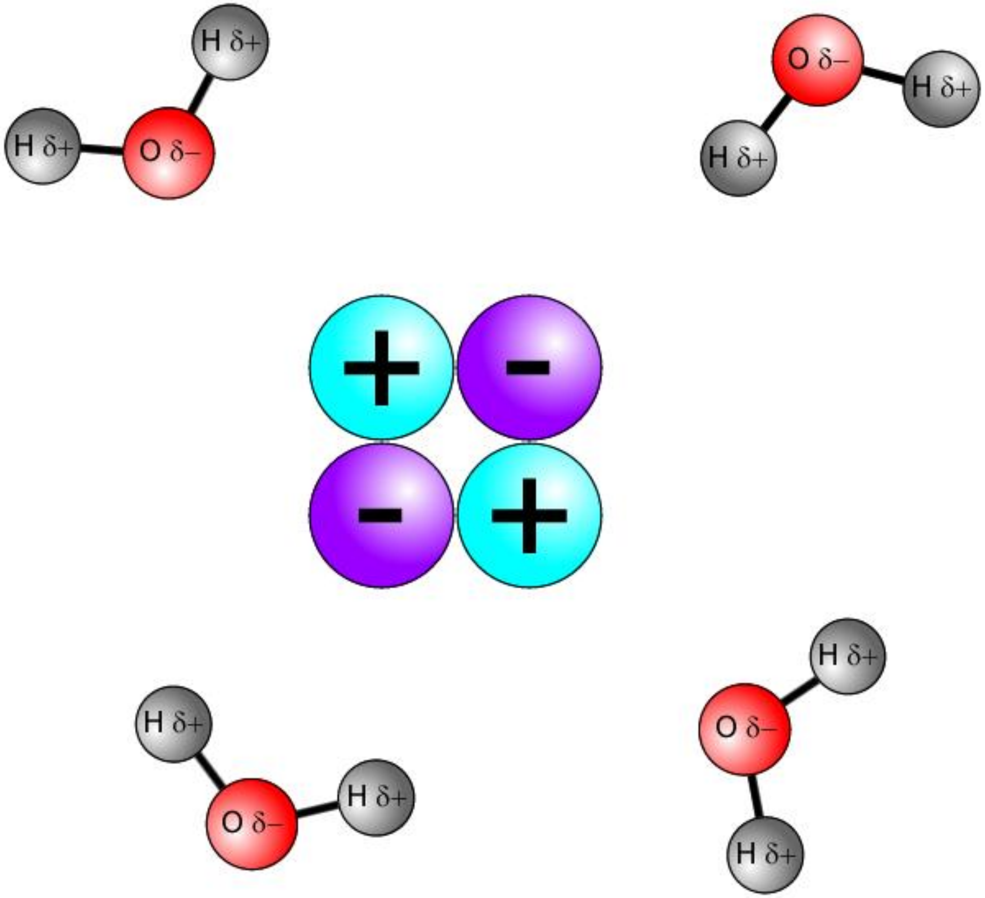
Energy from Chemicals – Enthalpy of Solution

- Dissolving an ionic compound in water can be considered to take place in *two stages*.

1. The electrostatic force of attraction between the oppositely charged anions and cations of the solute must be overcome. This process of *bond breaking* will *absorb energy* from the surroundings and is therefore *endothermic*, $+\Delta H$. Note: This enthalpy change is sometimes referred to as the *lattice enthalpy*.

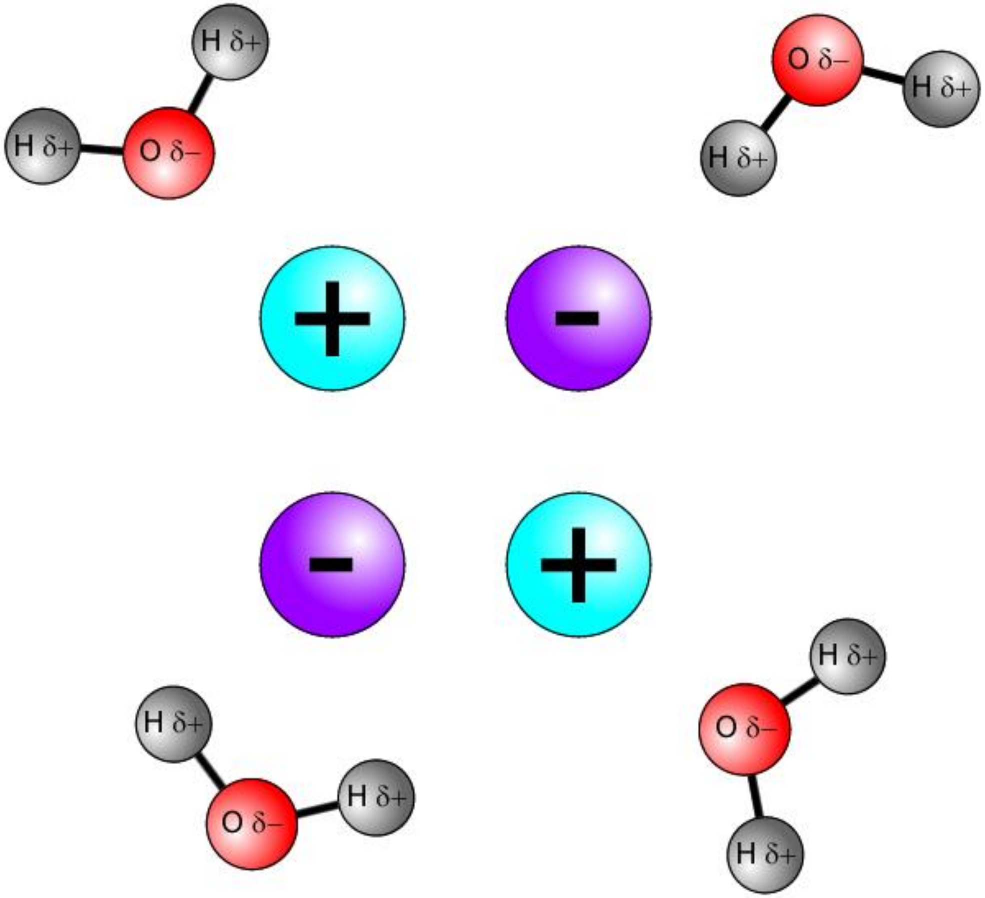
2. New bonds are formed between the anions and cations of the solute and the water molecules (solvent). This process of *bond formation* will *release energy* into the surroundings and is therefore *exothermic*, $-\Delta H$. Note: This enthalpy change is sometimes referred to as the *hydration enthalpy*.

Energy from Chemicals – Enthalpy of Solution



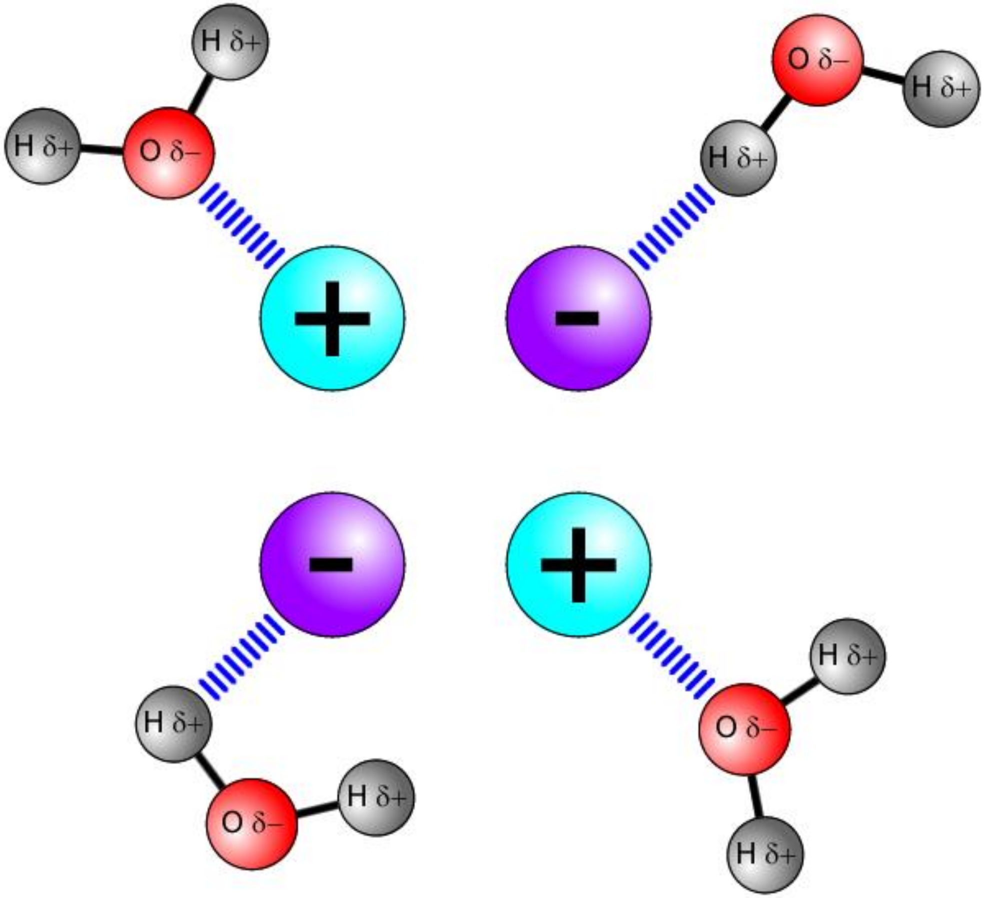
- The ionic compound (solute) is added to water (solvent).

Energy from Chemicals – Enthalpy of Solution



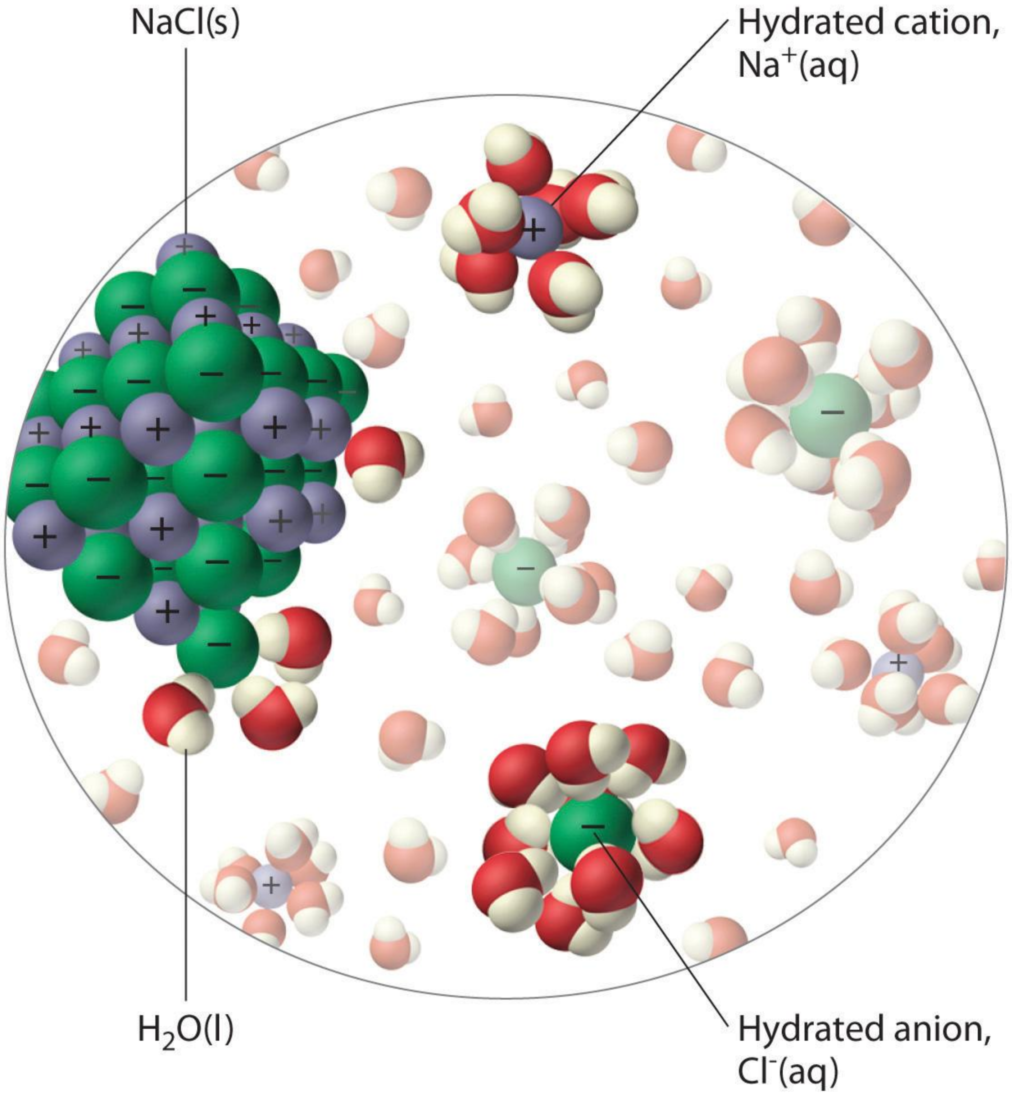
1. Energy is *absorbed* from the surroundings to overcome the electrostatic force of attraction between the anions and cations (*endothermic*, +ΔH).

Energy from Chemicals – Enthalpy of Solution



2. Energy is *released* into the surroundings when bonds are formed between the anions and cations of the solute and water molecules (*exothermic*, $-\Delta H$).

Energy from Chemicals – Enthalpy of Solution



Energy from Chemicals – Enthalpy of Solution

- The enthalpy change of solution for NaOH(s) is -45.0 kJ mol^{-1}. Using ideas about bond breaking and bond formation, comment on the significance of this enthalpy change.

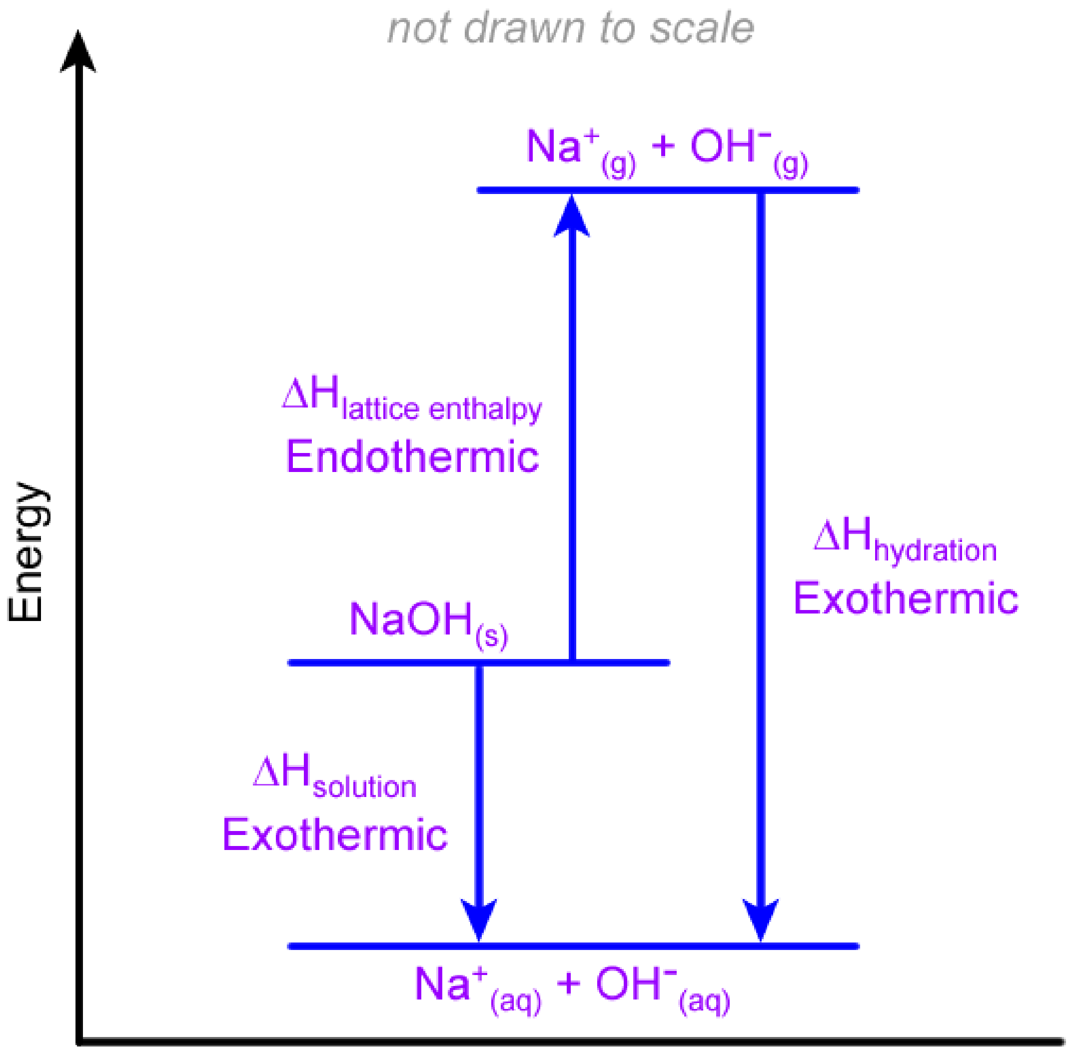
Energy from Chemicals – Enthalpy of Solution

- The enthalpy change of solution for NaOH(s) is –45.0 kJ mol^{-1}. Using ideas about bond breaking and bond formation, comment on the significance of this enthalpy change.

- The energy that is *released* into the surroundings during bond formation between Na^+ and H_2O, and OH^- and H_2O is *greater* than the energy that is *absorbed* from the surroundings to overcome the electrostatic force of attraction between Na^+ and OH^- of the solute. Overall, energy is *released* into the surroundings and the enthalpy change of solution for NaOH(s) is *exothermic*, $-\Delta H$.

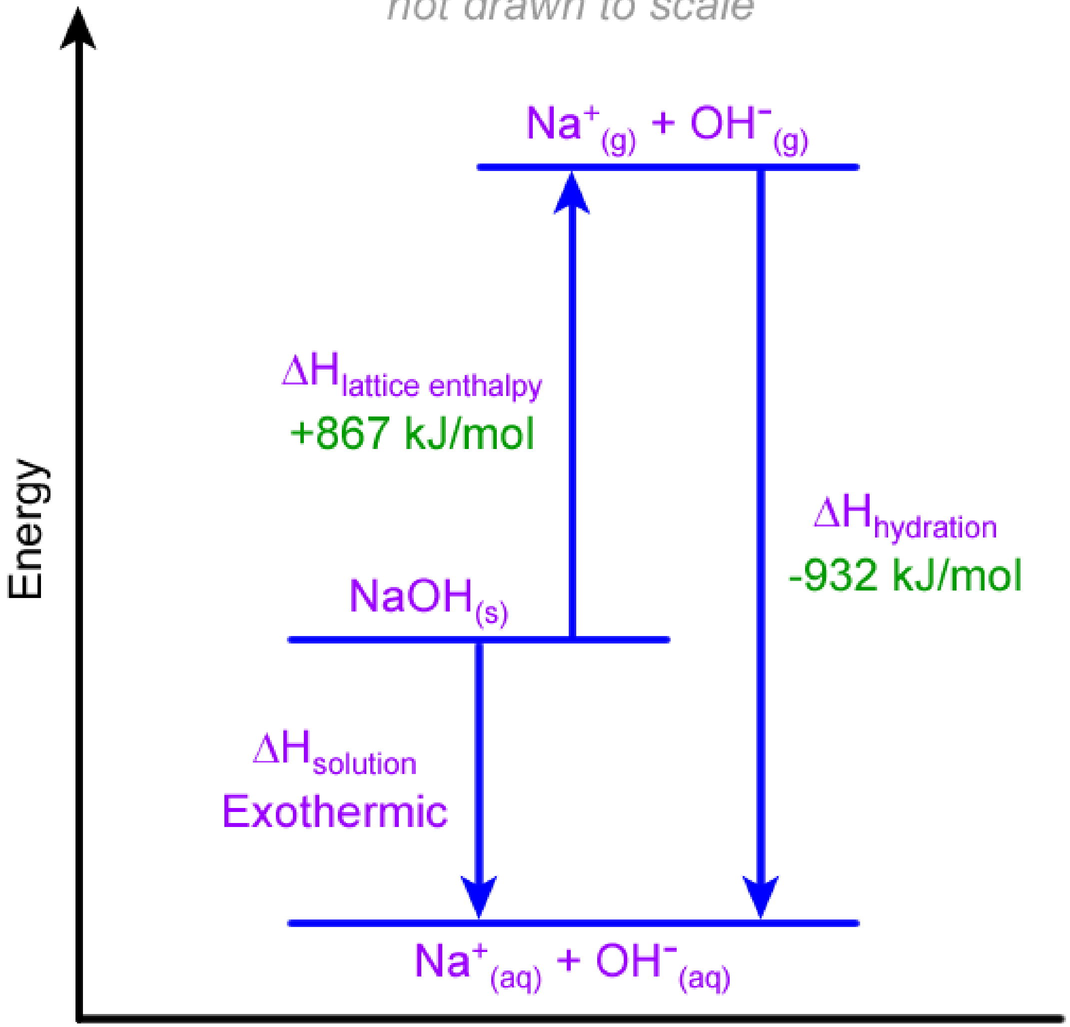
Energy from Chemicals – Enthalpy of Solution

energy level diagram



Energy from Chemicals – Enthalpy of Solution

energy level diagram



not drawn to scale

Energy

$Na^+_{(g)}$ + $OH^-_{(g)}$

$\Delta H_{\text{lattice enthalpy}}$

+867 kJ/mol

$\Delta H_{\text{hydration}}$

-932 kJ/mol

$NaOH_{(s)}$

$\Delta H_{\text{solution}}$

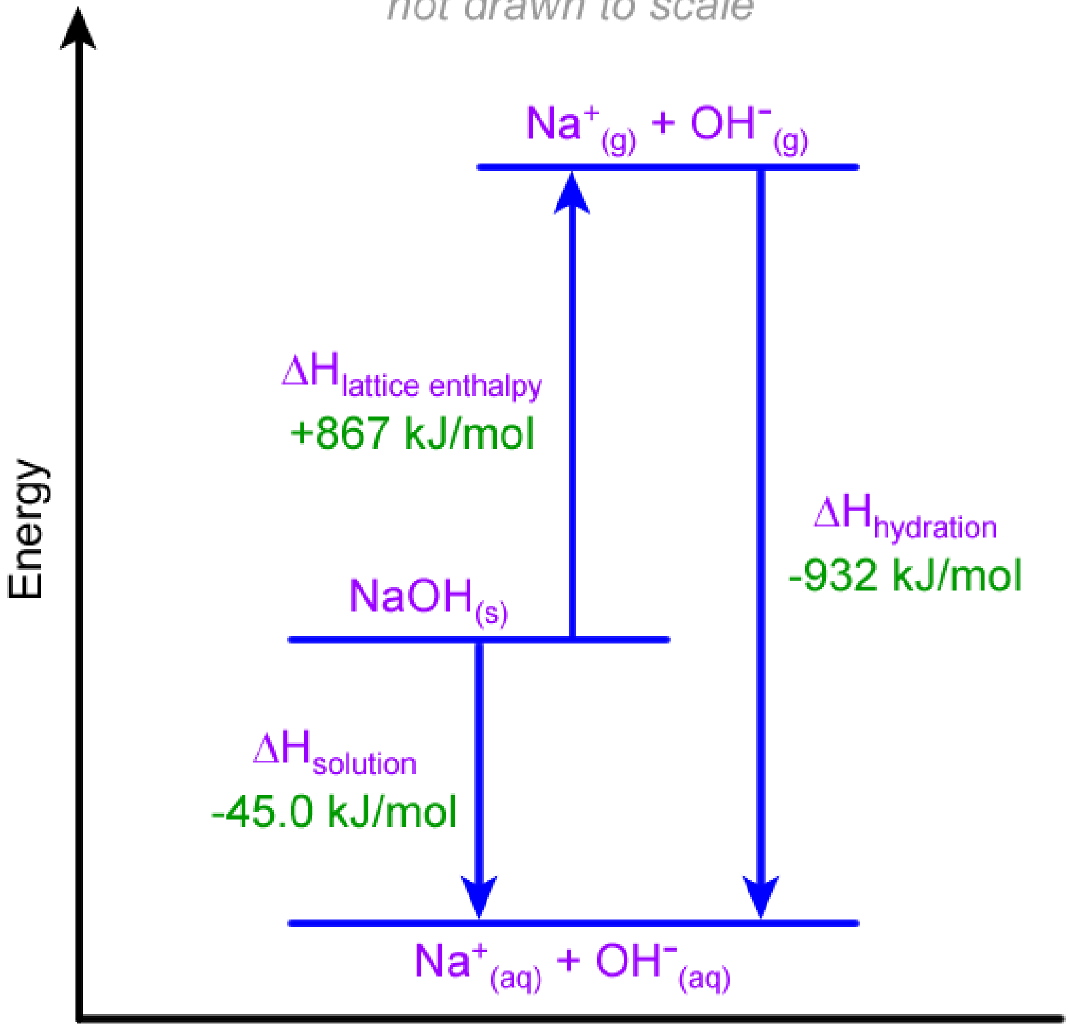
Exothermic

$Na^+_{(aq)}$ + $OH^-_{(aq)}$

Energy from Chemicals – Enthalpy of Solution

energy level diagram

not drawn to scale



$$\Delta H_{\text{solution}} = (+867) + (–932) = –45.0\ \text{kJ mol}^{–1}$$

Energy from Chemicals – Enthalpy of Solution

- The enthalpy change of solution for $NH_4NO_3(s)$ is $+25.0\ \text{kJ mol}^{-1}$. Using ideas about bond breaking and bond formation, comment on the significance of this enthalpy change.

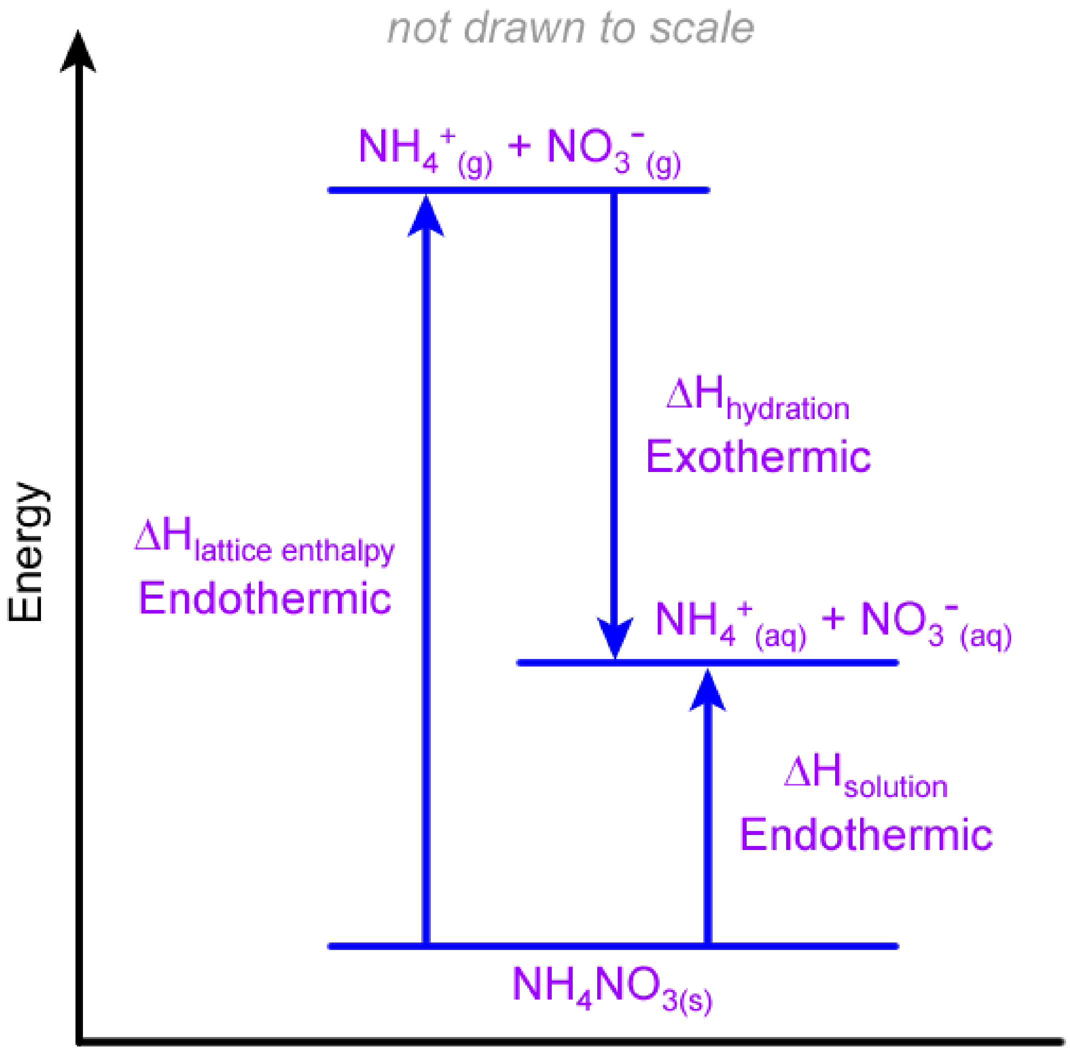
Energy from Chemicals – Enthalpy of Solution

- The enthalpy change of solution for NH_4NO_3(s) is +25.0 kJ mol^{-1}. Using ideas about bond breaking and bond formation, comment on the significance of this enthalpy change.

- The energy that is *absorbed* from the surroundings to overcome the electrostatic force of attraction between NH_4^+ and NO_3^- of the solute is *greater* than the energy that is *released* into the surroundings during bond formation between NH_4^+ and H_2O, and NO_3^- and H_2O. Overall, energy is *absorbed* from the surroundings and the enthalpy change of solution for NH_4NO_3(s) is *endothermic*, $+\Delta H$.

Energy from Chemicals – Enthalpy of Solution

energy level diagram



not drawn to scale

Energy

$NH_4{}^+{}_{(g)} + NO_3{}^-{}_{(g)}$

$\Delta H_{\text{lattice enthalpy}}$
Endothermic

$\Delta H_{\text{hydration}}$
Exothermic

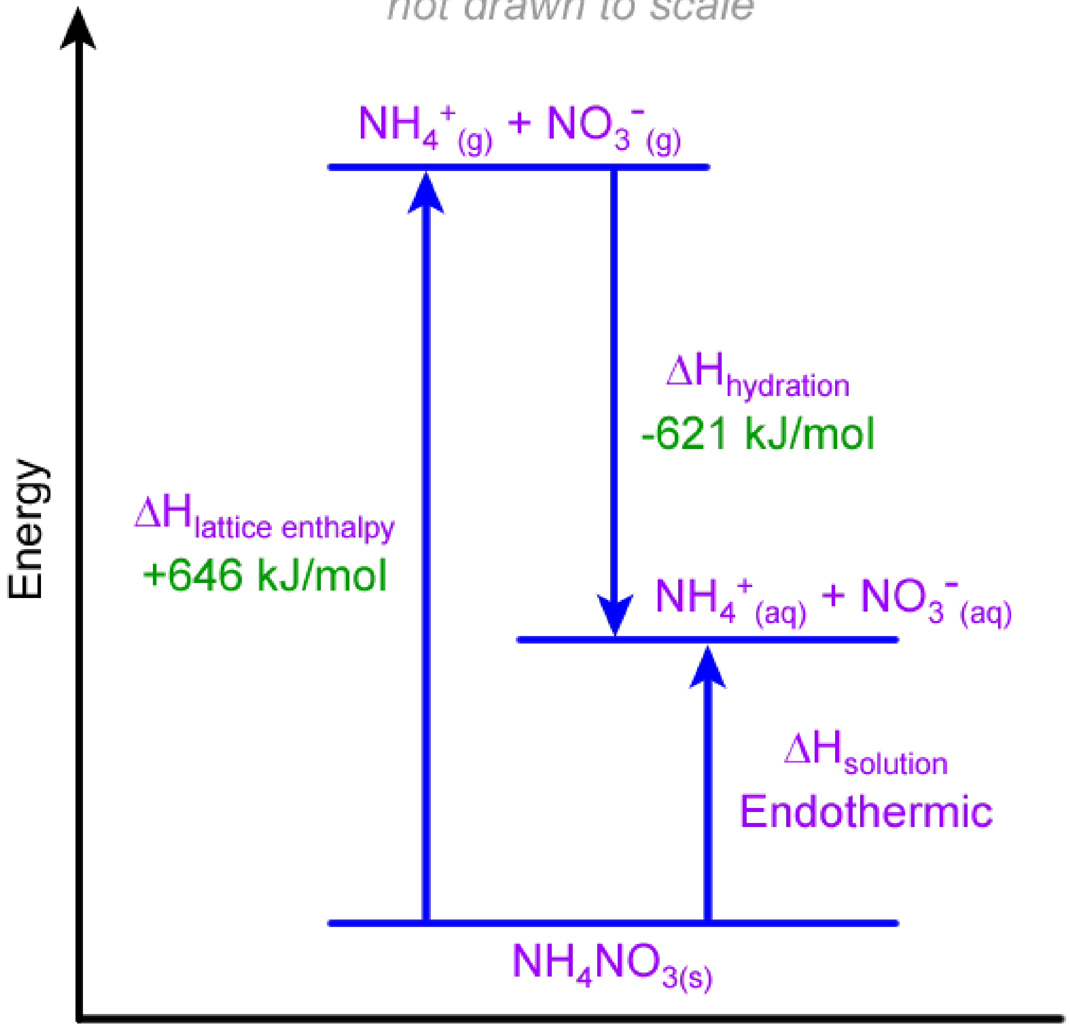
$NH_4{}^+{}_{(aq)} + NO_3{}^-{}_{(aq)}$

$\Delta H_{\text{solution}}$
Endothermic

$NH_4NO_{3(s)}$

Energy from Chemicals – Enthalpy of Solution

energy level diagram



not drawn to scale

Energy

$NH_4^+{}_{(g)} + NO_3^-{}_{(g)}$

$\Delta H_{\text{lattice enthalpy}}$

+646 kJ/mol

$\Delta H_{\text{hydration}}$

-621 kJ/mol

$NH_4^+{}_{(aq)} + NO_3^-{}_{(aq)}$

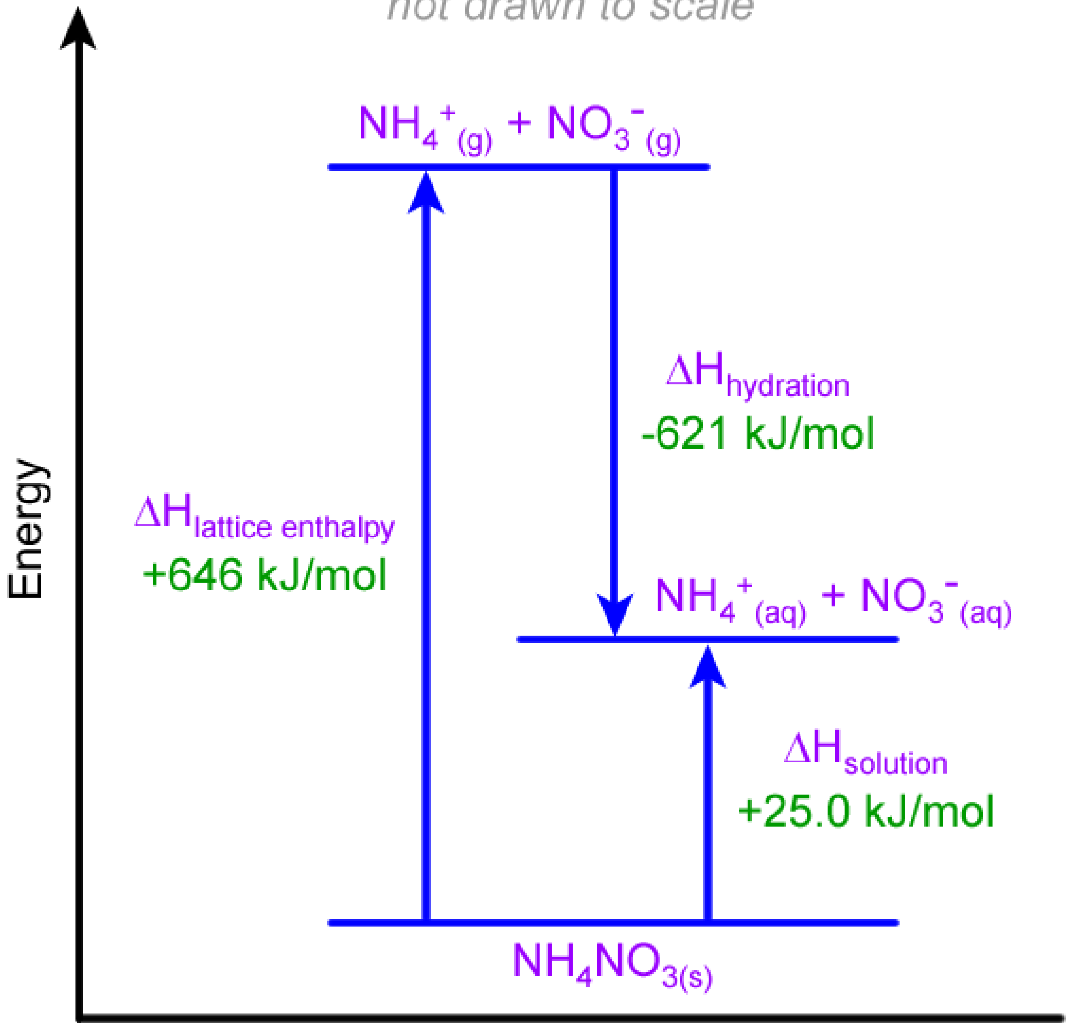
$\Delta H_{\text{solution}}$

Endothermic

$NH_4NO_{3(s)}$

Energy from Chemicals – Enthalpy of Solution

energy level diagram



not drawn to scale

Energy

$NH_4^+{}_{(g)} + NO_3^-{}_{(g)}$

$\Delta H_{lattice\ enthalpy}$
+646 kJ/mol

$\Delta H_{hydration}$
-621 kJ/mol

$NH_4^+{}_{(aq)} + NO_3^-{}_{(aq)}$

$\Delta H_{solution}$
+25.0 kJ/mol

$NH_4NO_{3(s)}$

$$\Delta H_{solution} = (+646) + (-621) = +25.0\ kJ\ mol^{-1}$$

Energy from Chemicals – Enthalpy of Solution

- The energy that must be *absorbed* from the surroundings in order to overcome the electrostatic force of attraction between an anion and a cation depends upon their *charge densities**. The *greater* the anion’s and cation’s charge densities, the *greater* the amount of energy that must be absorbed in order to overcome the electrostatic force of attraction between them.

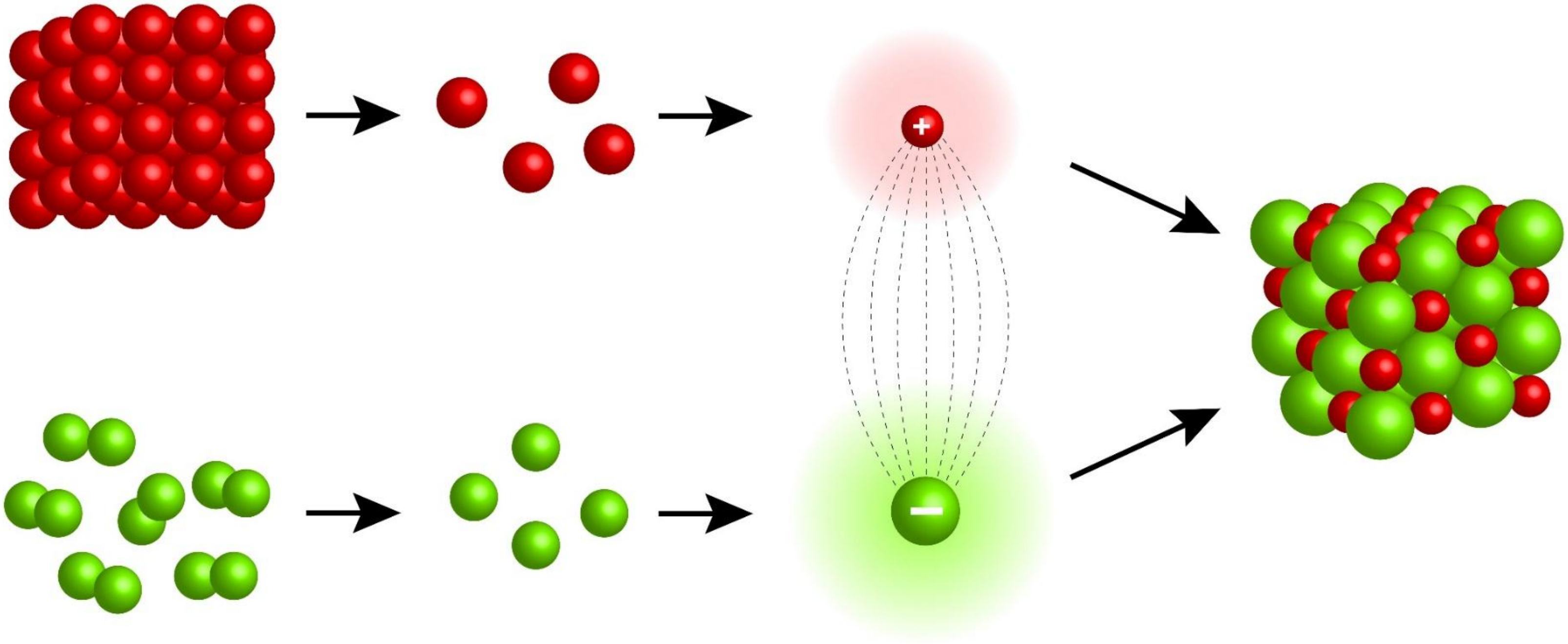
- The energy that is *released* into the surroundings when anions and cations are hydrated by water depends upon their *charge densities**. The *greater* an anion’s and cation’s charge densities, the *greater* the energy that will be released when water molecules bond to them.

*The *smaller* the ion’s *radius*, and the *greater* the ion’s *charge*, then the greater the ion’s *charge density*, *e.g.* Al^{3+} will have a relatively high charge density while Cs^+ will have a relatively small charge density.

Advance Concepts – Born-Haber Cycles

How do I calculate the *enthalpy change* of a reaction that has an *ionic product* (binary salt)?

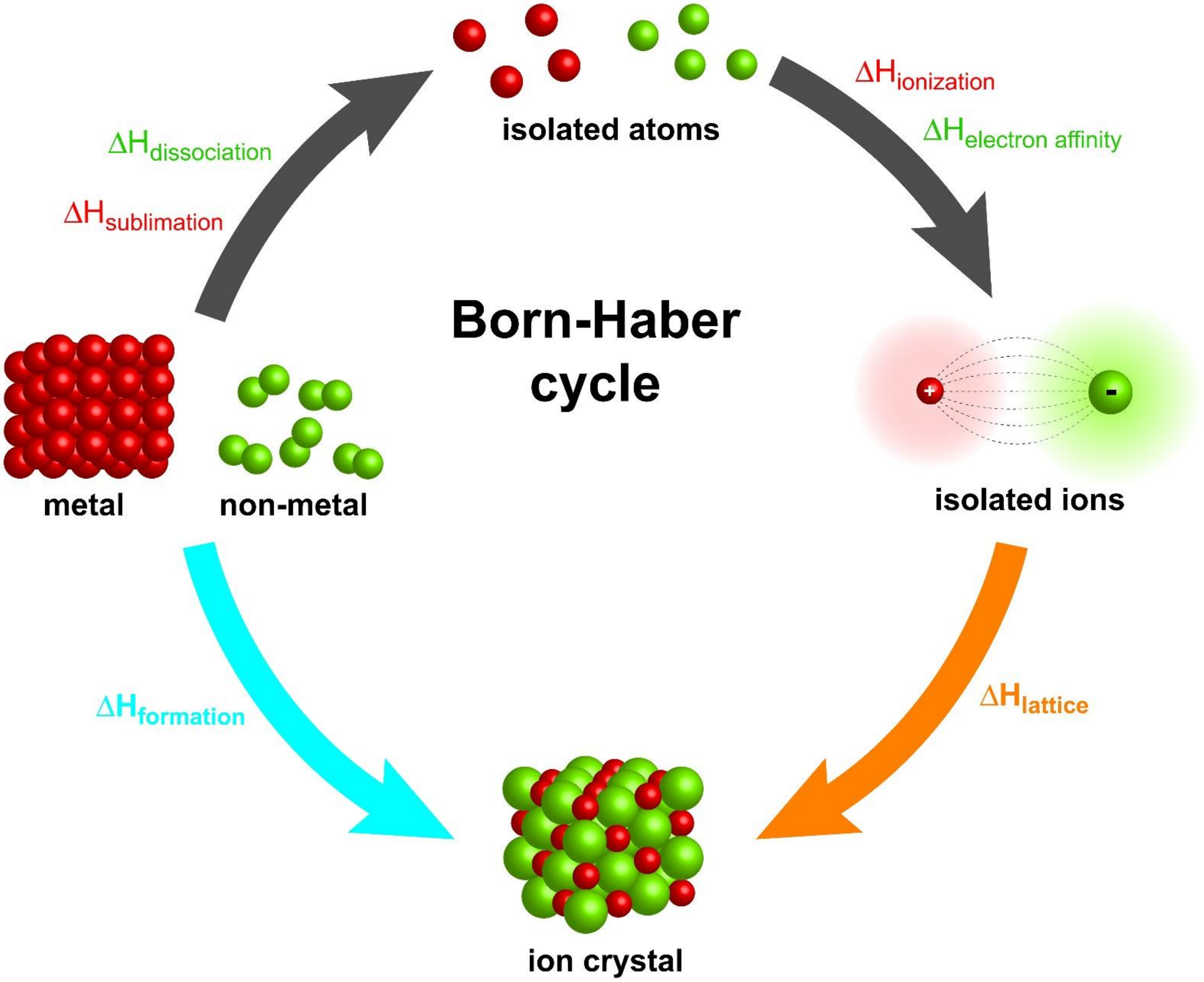
Advance Concepts – Born-Haber Cycles



An Overview of a Born-Haber Cycle

- The Born-Haber cycle takes account of all the enthalpy changes that occur when a metallic element and a non-metallic element react to form *one mole* of an ionic compound. The metallic and non-metallic elements must be *atomised* and *ionised* before they can combine together to form a crystal lattice.

Advance Concepts – Born-Haber Cycles



Born-Haber cycle

metal

non-metal

$\Delta H_{sublimation}$

$\Delta H_{dissociation}$

isolated atoms

$\Delta H_{ionization}$

$\Delta H_{electron\ affinity}$

isolated ions

$\Delta H_{lattice}$

ion crystal

$\Delta H_{formation}$

Advance Concepts – Born-Haber Cycles

Definitions of Terms Used

- Standard Enthalpy Change of Formation – $\Delta H^{\ominus}_{f}$:

The standard enthalpy change of formation is the enthalpy change when one mole of a compound is formed from its pure constituent elements, all substances being in their standard states at 1 atmosphere pressure (101.3 kPa) and 25°C (298 K).

Advance Concepts – Born-Haber Cycles

Definitions of Terms Used

- Standard Enthalpy of Atomisation – $\Delta H^{\ominus}_{at}$:

The standard enthalpy of atomisation is the enthalpy change when one mole of gaseous atoms is formed from its element in its defined physical state at 1 atmosphere pressure (101.3 kPa) and 25°C (298 K). The general equation for a diatomic element is:

$$^{1}/_{2}X_2(g) \rightarrow X(g)$$

Advance Concepts – Born-Haber Cycles

Definitions of Terms Used

- Bond Dissociation Enthalpy – $\Delta H^{\ominus}_{bde}$:

The bond dissociation enthalpy is the standard enthalpy change when one mole of chemical bonds are cleaved homolytically under standard conditions at 1 atmosphere pressure (101.3 kPa) and 25°C (298 K). The general equation for a diatomic element is:

$$X_2(g) \rightarrow 2X(g)$$

$$X - X(g) \rightarrow 2X(g)$$

Note: For a diatomic element, the value of the bond dissociation enthalpy is *twice* the value of the standard enthalpy of atomisation.

Advance Concepts – Born-Haber Cycles

Definitions of Terms Used

- Lattice Formation Enthalpy – $\Delta H^{\ominus}_{LA}$:

Lattice formation enthalpy is the standard enthalpy change when a solid ionic lattice is formed from its separate gaseous anions and cations. It is an exothermic process ($-\Delta H$). For example:

$$\mathrm{Na^+(g)} + \mathrm{C\mathit{l}^-(g)} \rightarrow \mathrm{NaC\mathit{l}(s)} \qquad \Delta H = -787\ \mathrm{kJ\ mol^{-1}}$$

Advance Concepts – Born-Haber Cycles

Definitions of Terms Used

- Lattice Dissociation Enthalpy – $\Delta H^{\ominus}_{LA}$:

Lattice dissociation enthalpy is the standard enthalpy change when a solid ionic lattice is dissociated into its component gaseous anions and cations which are separated by an infinite distance from each other. It is an endothermic process (+ΔH). For example:

$$\mathrm{NaC}l\mathrm{(s)} \rightarrow \mathrm{Na^{+}(g)} + \mathrm{C}l^{-}\mathrm{(g)} \quad \Delta H = +787\ \mathrm{kJ\ mol^{-1}}$$

Note: For the same ionic compound, the lattice formation enthalpy and lattice dissociation enthalpy have the same numerical value, but *opposite sign*.

Advance Concepts – Born-Haber Cycles

Definitions of Terms Used

- First Ionisation Energy – $\Delta H_{1st\ IE}$:

First ionisation energy is the energy absorbed (endothermic, $+\Delta H$) to convert one mole of gaseous atoms into one mole of gaseous cations, each cation carrying a charge of 1+. It is represented by the general equation:

$$X(g) \rightarrow X^{+}(g) + e^{-}$$

Advance Concepts – Born-Haber Cycles

Definitions of Terms Used

- Second Ionisation Energy – $\Delta H_{\text{2nd IE}}$:

Second ionisation energy is the energy absorbed (endothermic, +Δ) to convert one mole of gaseous cations, each carrying a charge of 1+, into one mole of gaseous cations, each carrying a charge of 2+. It is represented by the general equation:

$$X^{+}(g) \rightarrow X^{2+}(g) + e^{-}$$

Advance Concepts – Born-Haber Cycles

Definitions of Terms Used

- First Electron Affinity – $\Delta H_{1st\ ea}$:

First electron affinity is the energy released (exothermic, $-\Delta H$) when one mole of gaseous atoms is converted into one mole of gaseous anions, each anion carrying a charge of 1–. It is represented by the general equation:

$$X(g) \ + \ e^{-} \ \rightarrow \ X^{-}(g)$$

Advance Concepts – Born-Haber Cycles

Definitions of Terms Used

- Second Electron Affinity – $\Delta H_{2nd\ ea}$:

Second electron affinity is the energy absorbed (endothermic, $+\Delta H$) when one mole of gaseous anions, each carrying a charge of 1–, is converted into one mole of gaseous anions, each carrying a charge of 2–. It is represented by the general equation:

$$X^{-}(g) \; + \; e^{-} \; \rightarrow \; X^{2-}(g)$$

Advance Concepts – Born-Haber Cycles

- The following data can be used to construct a Born-Haber cycle to represent the formation of one mole of sodium chloride from the elements sodium and chlorine in their standard states.

$\mathrm{Na(s)} + {}^{1}/_{2}\mathrm{C}l_{2}\mathrm{(g)} \rightarrow \mathrm{NaC}l\mathrm{(s)} \quad \Delta H_{formation} = -410\ \mathrm{kJ\ mol^{-1}}$

$\mathrm{Na(s)} \rightarrow \mathrm{Na(g)} \quad \Delta H_{atomisation} = +108\ \mathrm{kJ\ mol^{-1}}$

${}^{1}/_{2}\mathrm{C}l_{2}\mathrm{(g)} \rightarrow \mathrm{C}l\mathrm{(g)} \quad \Delta H_{atomisation} = +122\ \mathrm{kJ\ mol^{-1}}$

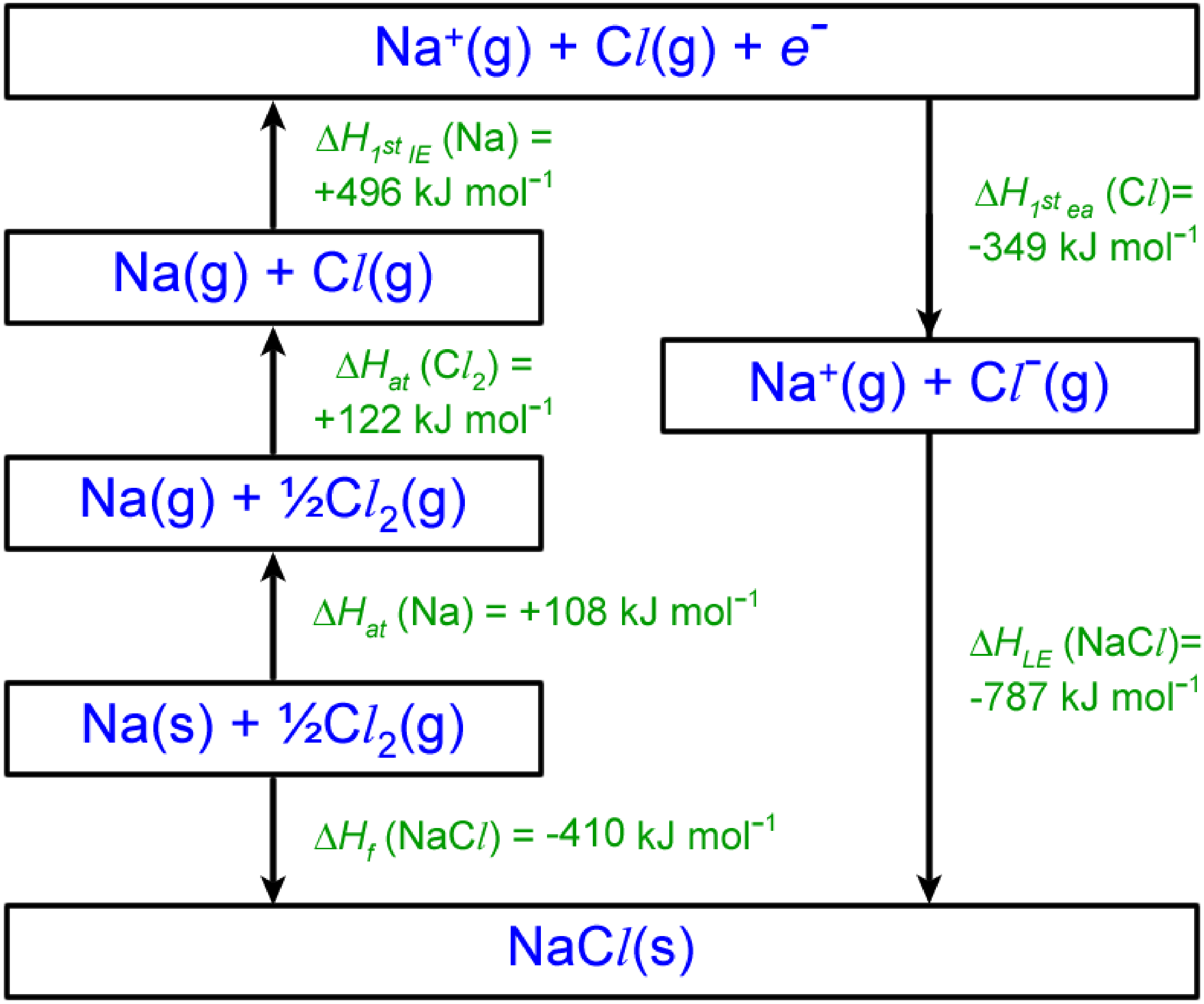
$\mathrm{Na(g)} \rightarrow \mathrm{Na^{+}(g)} + e^{-} \quad \Delta H_{first\ ionisation} = +496\ \mathrm{kJ\ mol^{-1}}$

$\mathrm{C}l\mathrm{(g)} + e^{-} \rightarrow \mathrm{C}l^{-}\mathrm{(g)} \quad \Delta H_{first\ electron\ affinity} = -349\ \mathrm{kJ\ mol^{-1}}$

$\mathrm{Na^{+}(g)} + \mathrm{C}l^{-}\mathrm{(g)} \rightarrow \mathrm{NaC}l\mathrm{(s)} \quad \Delta H_{lattice\ enthalpy} = -787\ \mathrm{kJ\ mol^{-1}}$

Advance Concepts – Born-Haber Cycles

Born-Haber Cycle for Sodium Chloride



$\mathrm{Na^{+}(g) + C}l\mathrm{(g)} + e^{-}$

$\Delta H_{1st\ IE}$ (Na) = +496 kJ mol^{-1}

$\Delta H_{1st\ ea}$ (Cl)= -349 kJ mol^{-1}

$\mathrm{Na(g) + C}l\mathrm{(g)}$

$\mathrm{Na^{+}(g) + C}l^{-}\mathrm{(g)}$

ΔH_{at} (Cl_2) = +122 kJ mol^{-1}

$\mathrm{Na(g) + ½C}l_2\mathrm{(g)}$

ΔH_{at} (Na) = +108 kJ mol^{-1}

ΔH_{LE} (NaCl)= -787 kJ mol^{-1}

$\mathrm{Na(s) + ½C}l_2\mathrm{(g)}$

ΔH_f (NaCl) = -410 kJ mol^{-1}

$\mathrm{NaC}l\mathrm{(s)}$

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Advance Concepts – Born-Haber Cycles

- Use the data given below to calculate the *enthalpy change of formation* of calcium fluoride from the elements calcium and fluorine in their standard states.

$\mathrm{Ca(s)} + \mathrm{F_2(g)} \rightarrow \mathrm{CaF_2(s)} \quad \Delta H_{formation} = \text{unknown kJ mol}^{-1}$

$\mathrm{Ca(s)} \rightarrow \mathrm{Ca(g)} \quad \Delta H_{atomisation} = +179 \text{ kJ mol}^{-1}$

$\mathrm{F_2(g)} \rightarrow 2\mathrm{F(g)} \quad \Delta H_{atomisation} = +69.5 \text{ kJ mol}^{-1}$

$\mathrm{Ca(g)} \rightarrow \mathrm{Ca^{+}(g)} + e^{-} \quad \Delta H_{first\ ionisation\ energy} = +590 \text{ kJ mol}^{-1}$

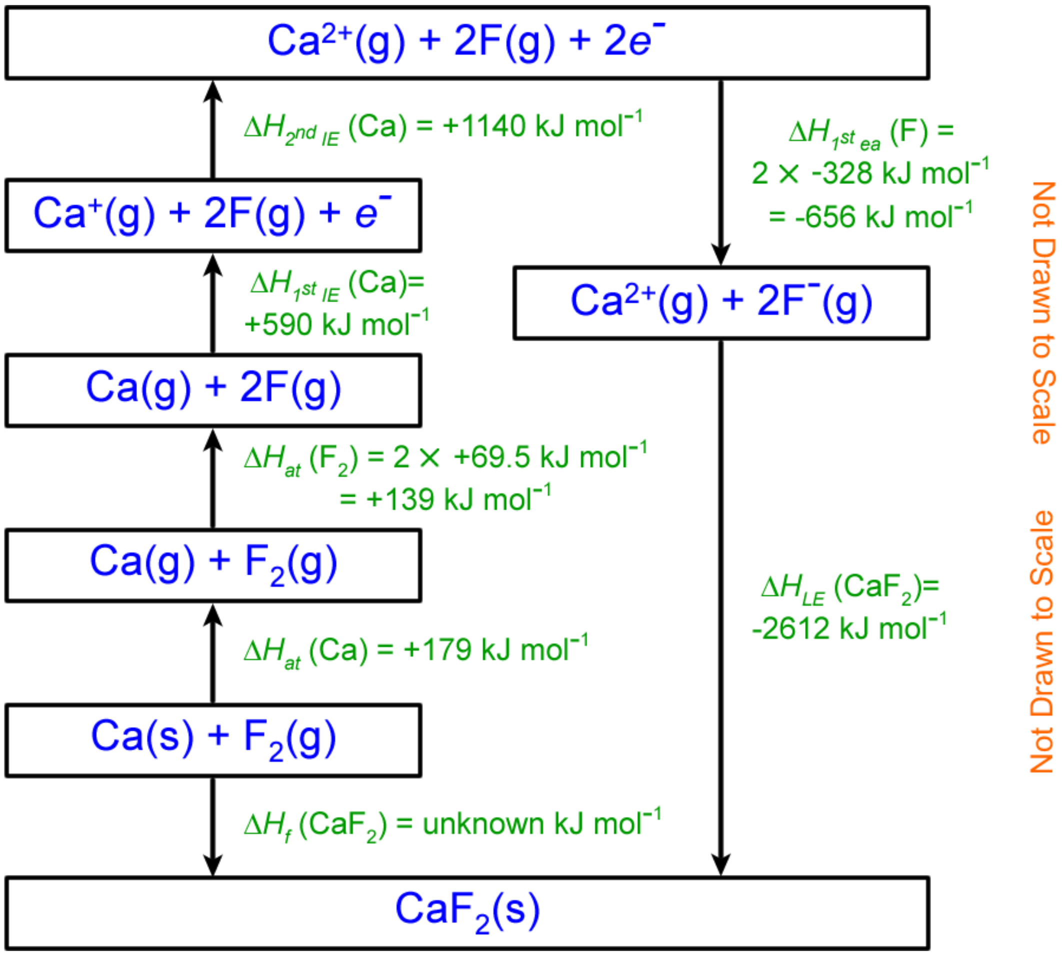
$\mathrm{Ca^{+}(g)} \rightarrow \mathrm{Ca^{2+}(g)} + e^{-} \quad \Delta H_{second\ ionisation\ energy} = +1140 \text{ kJ mol}^{-1}$

$2\mathrm{F(g)} + 2e^{-} \rightarrow 2\mathrm{F^{-}(g)} \quad \Delta H_{first\ electron\ affinity} = -656 \text{ kJ mol}^{-1}$

$\mathrm{Ca^{2+}(g)} + 2\mathrm{F^{-}(g)} \rightarrow \mathrm{CaF_2(s)} \quad \Delta H_{lattice\ enthalpy} = -2612 \text{ kJ mol}^{-1}$

Advance Concepts – Born-Haber Cycles

Born-Haber Cycle to Calculate ΔH_f for CaF_2



$Ca^{2+}(g) + 2F(g) + 2e^-$

$\Delta H_{2nd\ IE}$ (Ca) = +1140 kJ mol^{-1}

$Ca^{+}(g) + 2F(g) + e^-$

$\Delta H_{1st\ IE}$ (Ca)= +590 kJ mol^{-1}

$Ca(g) + 2F(g)$

ΔH_{at} (F_2) = 2 × +69.5 kJ mol^{-1} = +139 kJ mol^{-1}

$Ca(g) + F_2(g)$

ΔH_{at} (Ca) = +179 kJ mol^{-1}

$Ca(s) + F_2(g)$

ΔH_f (CaF_2) = unknown kJ mol^{-1}

$CaF_2(s)$

$\Delta H_{1st\ ea}$ (F) = 2 × -328 kJ mol^{-1} = -656 kJ mol^{-1}

$Ca^{2+}(g) + 2F^-(g)$

ΔH_{LE} (CaF_2)= -2612 kJ mol^{-1}

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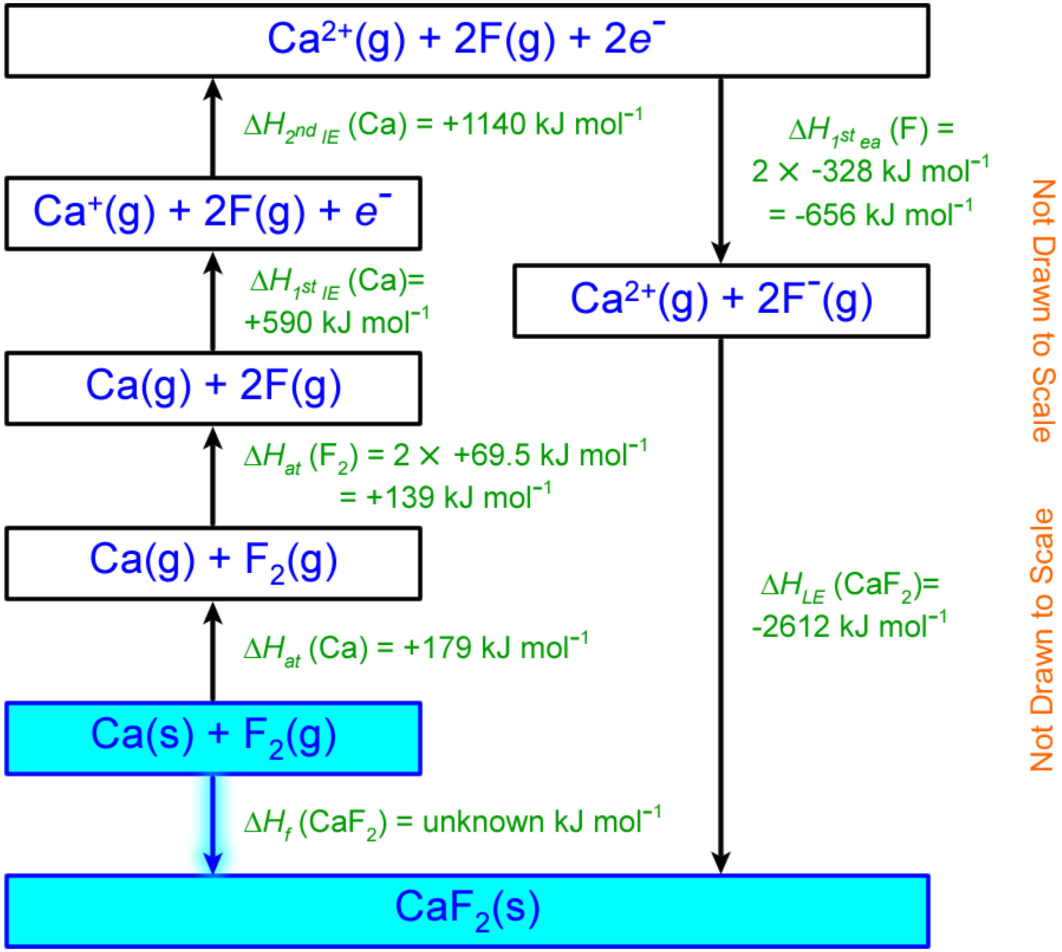
- Calculate the enthalpy change of formation for CaF_2(s).

Advance Concepts – Born-Haber Cycles

Born-Haber Cycle to Calculate ΔH_f for CaF_2



$Ca^{2+}(g) + 2F(g) + 2e^-$

$\Delta H_{2nd\ IE}$ (Ca) = +1140 kJ mol^{-1}

$Ca^+(g) + 2F(g) + e^-$

$\Delta H_{1st\ IE}$ (Ca)= +590 kJ mol^{-1}

$Ca(g) + 2F(g)$

ΔH_{at} (F_2) = 2 × +69.5 kJ mol^{-1} = +139 kJ mol^{-1}

$Ca(g) + F_2(g)$

ΔH_{at} (Ca) = +179 kJ mol^{-1}

$Ca(s) + F_2(g)$

ΔH_f (CaF_2) = unknown kJ mol^{-1}

$\Delta H_{1st\ ea}$ (F) = 2 × -328 kJ mol^{-1} = -656 kJ mol^{-1}

$Ca^{2+}(g) + 2F^-(g)$

ΔH_{LE} (CaF_2)= -2612 kJ mol^{-1}

$CaF_2(s)$

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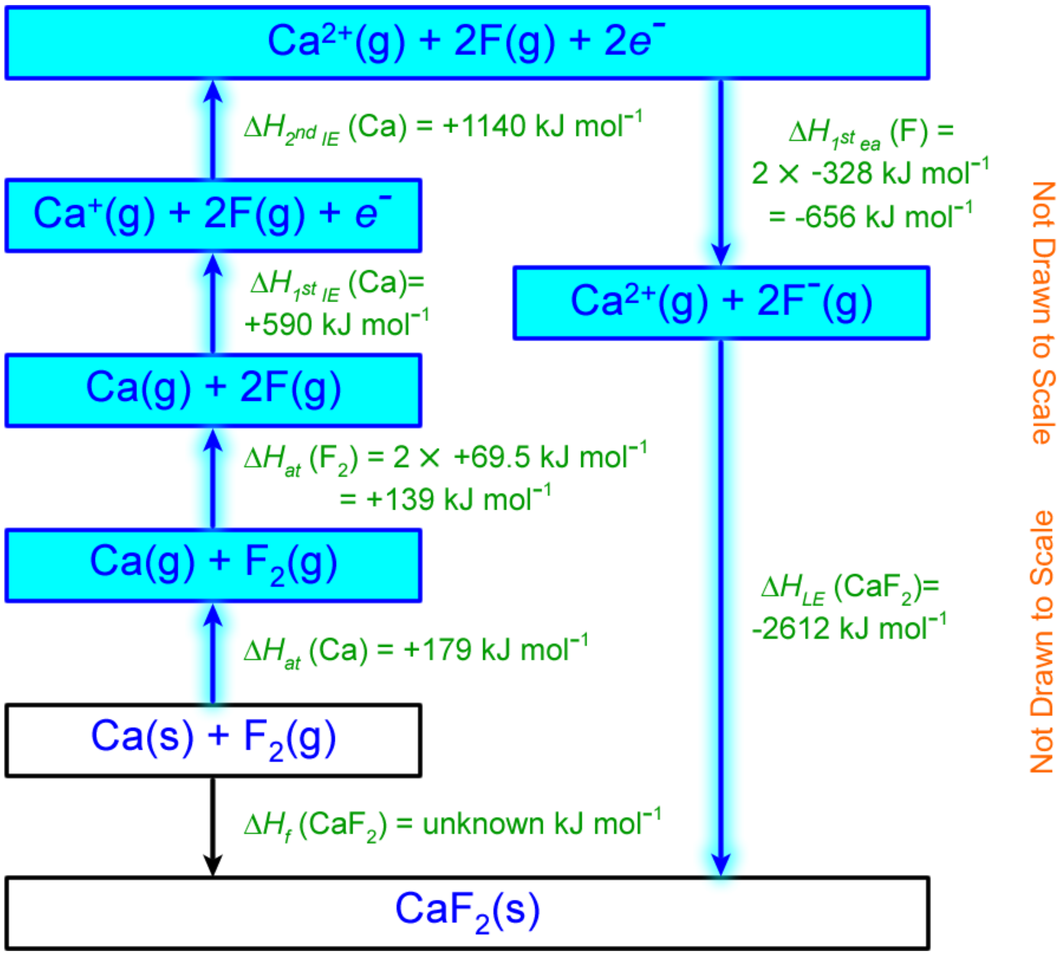
- It is not possible to go directly from $Ca(s) + F_2(g)$ to $CaF_2(s)$

Advance Concepts – Born-Haber Cycles

Born-Haber Cycle to Calculate ΔH_f for CaF_2



$Ca^{2+}(g) + 2F(g) + 2e^-$

$\Delta H_{2nd\ IE}$ (Ca) = +1140 kJ mol^{-1}

$\Delta H_{1st\ ea}$ (F) = 2 × -328 kJ mol^{-1} = -656 kJ mol^{-1}

$Ca^{+}(g) + 2F(g) + e^-$

$\Delta H_{1st\ IE}$ (Ca)= +590 kJ mol^{-1}

$Ca^{2+}(g) + 2F^-(g)$

$Ca(g) + 2F(g)$

ΔH_{at} (F_2) = 2 × +69.5 kJ mol^{-1} = +139 kJ mol^{-1}

$Ca(g) + F_2(g)$

ΔH_{LE} (CaF_2)= -2612 kJ mol^{-1}

ΔH_{at} (Ca) = +179 kJ mol^{-1}

$Ca(s) + F_2(g)$

ΔH_f (CaF_2) = unknown kJ mol^{-1}

$CaF_2(s)$

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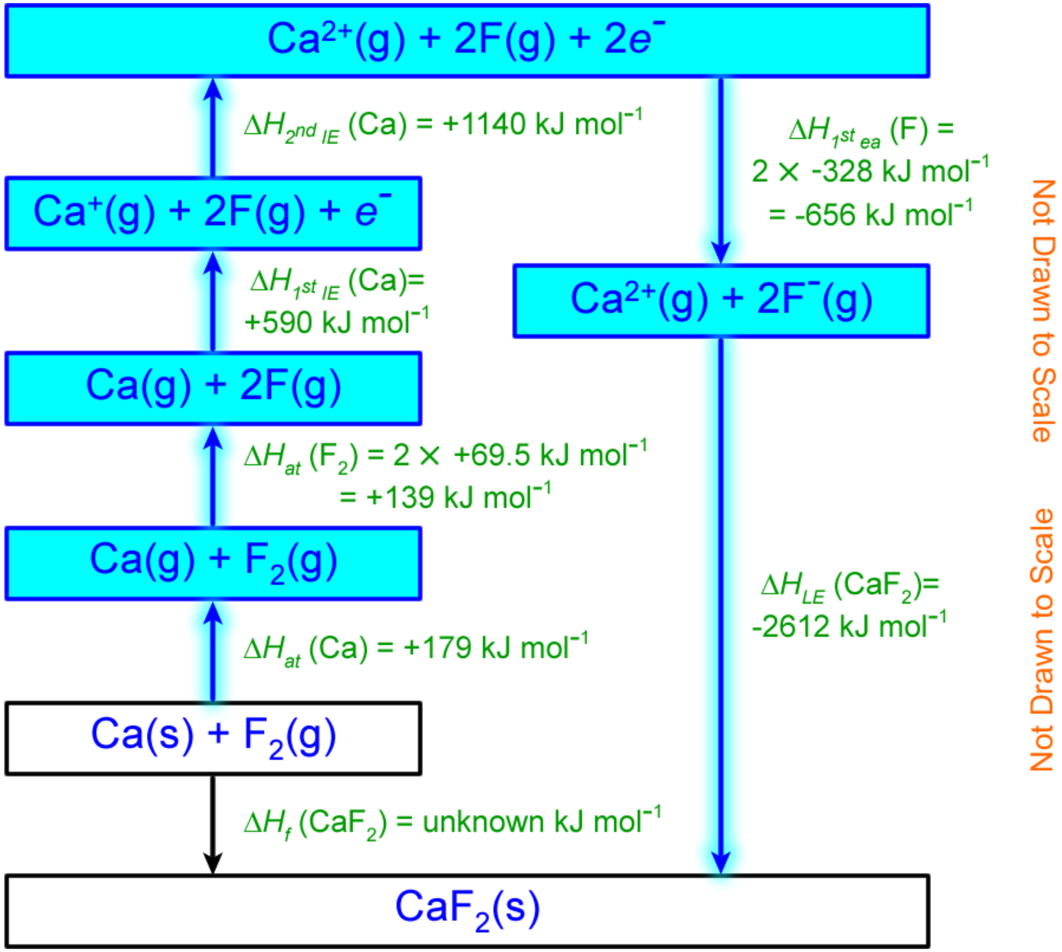
- Follow the direction of the blue arrows through the cycle.
- The sum of the enthalpy changes will give the ΔH_f for CaF_2.

Advance Concepts – Born-Haber Cycles

Born-Haber Cycle to Calculate ΔH_f for CaF_2



$Ca^{2+}(g) + 2F(g) + 2e^-$

$\Delta H_{2nd\ IE}$ (Ca) = +1140 kJ mol^{-1}

$\Delta H_{1st\ ea}$ (F) = 2 × -328 kJ mol^{-1} = -656 kJ mol^{-1}

$Ca^{+}(g) + 2F(g) + e^-$

$\Delta H_{1st\ IE}$ (Ca)= +590 kJ mol^{-1}

$Ca^{2+}(g) + 2F^-(g)$

$Ca(g) + 2F(g)$

ΔH_{at} (F_2) = 2 × +69.5 kJ mol^{-1} = +139 kJ mol^{-1}

$Ca(g) + F_2(g)$

ΔH_{LE} (CaF_2)= -2612 kJ mol^{-1}

ΔH_{at} (Ca) = +179 kJ mol^{-1}

$Ca(s) + F_2(g)$

ΔH_f (CaF_2) = unknown kJ mol^{-1}

$CaF_2(s)$

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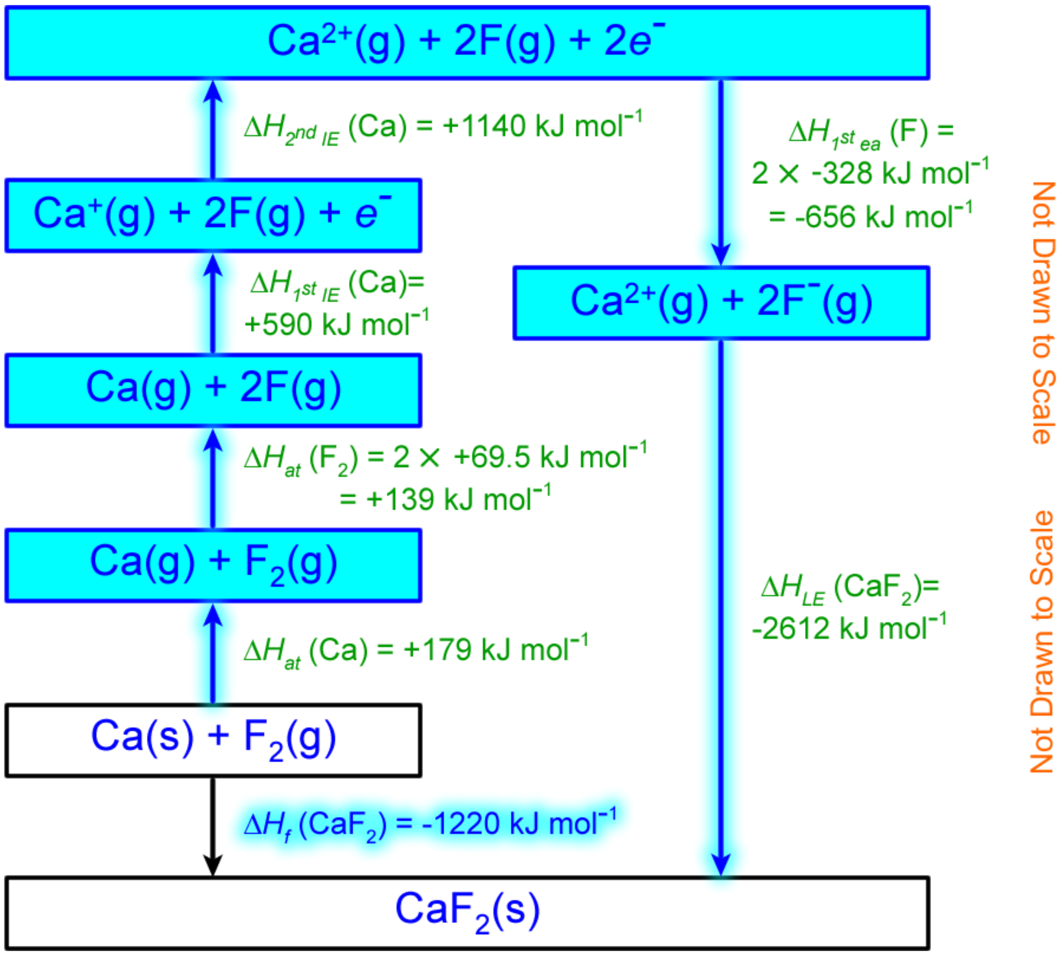
$$\Delta H_f (CaF_2) = (+179) + (+139) + (+590) + (+1140) + (–656) + (–2612)$$

Advance Concepts – Born-Haber Cycles

Born-Haber Cycle to Calculate ΔH_f for CaF_2



$Ca^{2+}(g) + 2F(g) + 2e^-$

$\Delta H_{2nd\ IE}$ (Ca) = +1140 kJ mol^{-1}

$Ca^{+}(g) + 2F(g) + e^-$

$\Delta H_{1st\ IE}$ (Ca)= +590 kJ mol^{-1}

$Ca(g) + 2F(g)$

ΔH_{at} (F_2) = 2 × +69.5 kJ mol^{-1} = +139 kJ mol^{-1}

$Ca(g) + F_2(g)$

ΔH_{at} (Ca) = +179 kJ mol^{-1}

$Ca(s) + F_2(g)$

ΔH_f (CaF_2) = -1220 kJ mol^{-1}

$\Delta H_{1st\ ea}$ (F) = 2 × -328 kJ mol^{-1} = -656 kJ mol^{-1}

$Ca^{2+}(g) + 2F^-(g)$

ΔH_{LE} (CaF_2)= -2612 kJ mol^{-1}

$CaF_2(s)$

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$$\Delta H_f (CaF_2) = (+179) + (+139) + (+590) + (+1140) + (–656) + (–2612)$$

$$\Delta H_f (CaF_2) = –1220 \text{ kJ mol}^{–1}$$

Advance Concepts – Born-Haber Cycles

- Use the data given below to calculate the *lattice formation enthalpy* of sodium oxide.

$2Na(s) + {}^{1}/_{2}O_2(g) \rightarrow Na_2O(s) \quad \Delta H_{formation} = +414\ kJ\ mol^{-1}$

$2Na(s) \rightarrow 2Na(g) \quad \Delta H_{atomisation} = +216\ kJ\ mol^{-1}$

${}^{1}/_{2}O_2(g) \rightarrow O(g) \quad \Delta H_{atomisation} = +249\ kJ\ mol^{-1}$

$Na(g) \rightarrow Na^+(g) + e^- \quad \Delta H_{first\ ionisation} = +992\ kJ\ mol^{-1}$

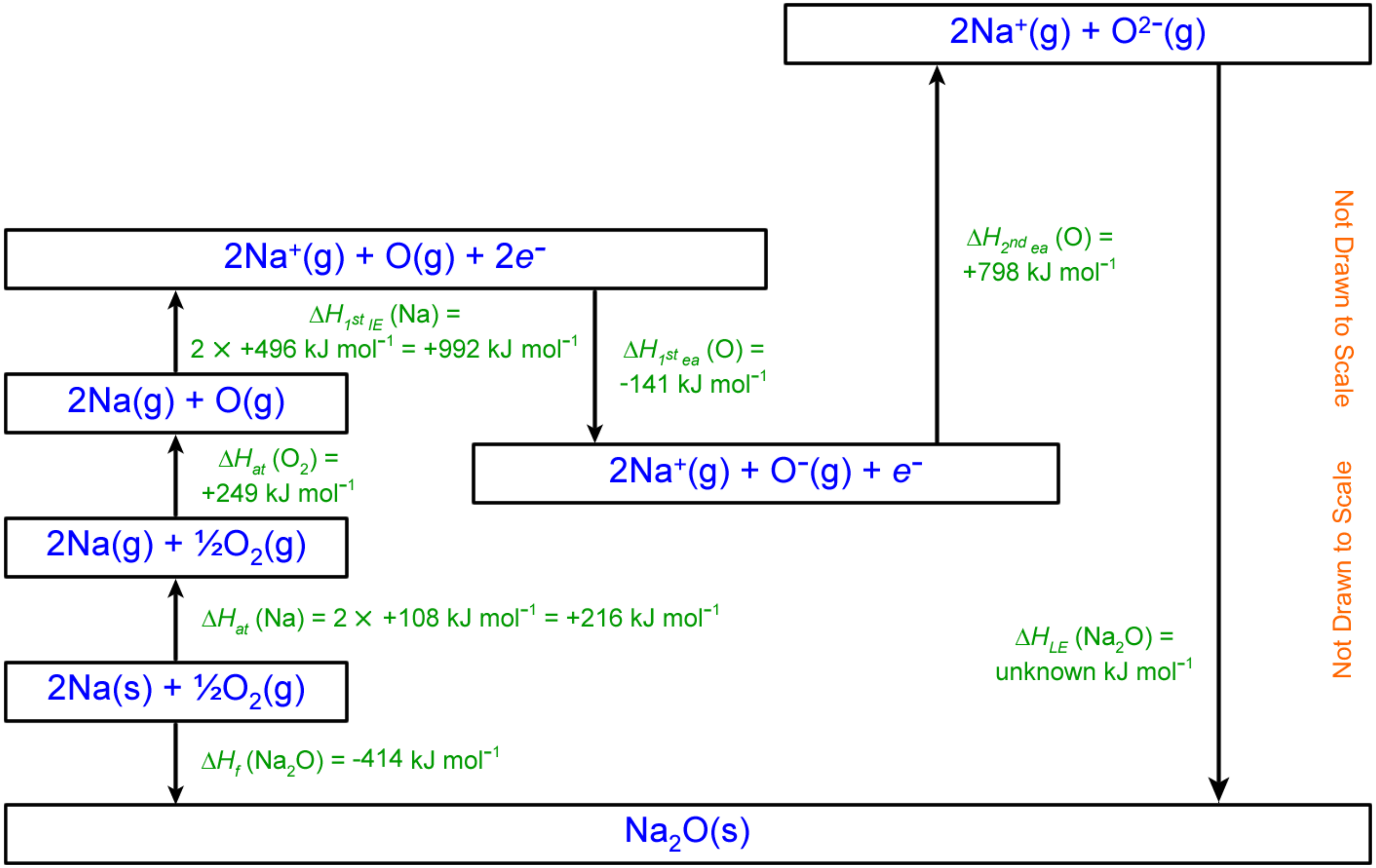
$O(g) + e^- \rightarrow O^-(g) \quad \Delta H_{first\ electron\ affinity} = -141\ kJ\ mol^{-1}$

$O^-(g) + e^- \rightarrow O^{2-}(g) \quad \Delta H_{second\ electron\ affinity} = +798\ kJ\ mol^{-1}$

$2Na^+(g) + O^{2-}(g) \rightarrow Na_2O(s) \quad \Delta H_{lattice\ enthalpy} = \text{unknown}\ kJ\ mol^{-1}$

Advance Concepts – Born-Haber Cycles

Born-Haber Cycle to Calculate ΔH$_{LE}$ for Na_2O



$2Na^+(g) + O^{2-}(g)$

$\Delta H_{2nd\ ea}$ (O) = +798 kJ mol^{-1}

$2Na^+(g) + O(g) + 2e^-$

$\Delta H_{1st\ IE}$ (Na) = 2 × +496 kJ mol^{-1} = +992 kJ mol^{-1}

$\Delta H_{1st\ ea}$ (O) = -141 kJ mol^{-1}

$2Na(g) + O(g)$

$2Na^+(g) + O^-(g) + e^-$

ΔH_{at} (O_2) = +249 kJ mol^{-1}

$2Na(g) + \frac{1}{2}O_2(g)$

ΔH_{at} (Na) = 2 × +108 kJ mol^{-1} = +216 kJ mol^{-1}

ΔH_{LE} (Na_2O) = unknown kJ mol^{-1}

$2Na(s) + \frac{1}{2}O_2(g)$

ΔH_f (Na_2O) = -414 kJ mol^{-1}

$Na_2O(s)$

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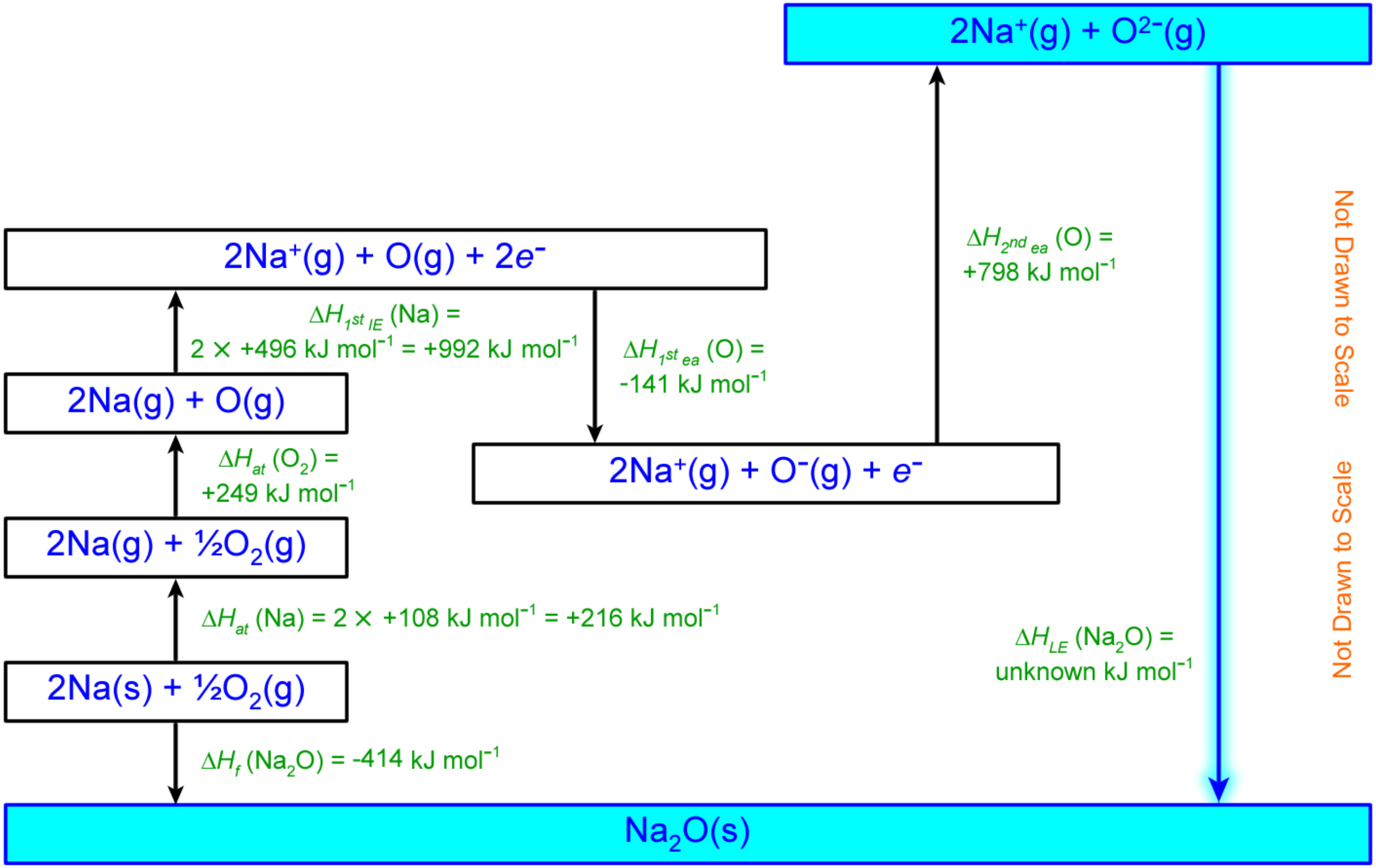
- Calculate the lattice formation enthalpy for $Na_2O(s)$.

Advance Concepts – Born-Haber Cycles

Born-Haber Cycle to Calculate ΔH_{LE} for Na_2O



$2Na^+(g) + O^{2-}(g)$

$\Delta H_{2nd\ ea}$ (O) = +798 kJ mol^{-1}

$2Na^+(g) + O(g) + 2e^-$

$\Delta H_{1st\ IE}$ (Na) = 2 × +496 kJ mol^{-1} = +992 kJ mol^{-1}

$\Delta H_{1st\ ea}$ (O) = -141 kJ mol^{-1}

$2Na(g) + O(g)$

$2Na^+(g) + O^-(g) + e^-$

ΔH_{at} (O_2) = +249 kJ mol^{-1}

$2Na(g) + ½O_2(g)$

ΔH_{at} (Na) = 2 × +108 kJ mol^{-1} = +216 kJ mol^{-1}

ΔH_{LE} (Na_2O) = unknown kJ mol^{-1}

$2Na(s) + ½O_2(g)$

ΔH_f (Na_2O) = -414 kJ mol^{-1}

$Na_2O(s)$

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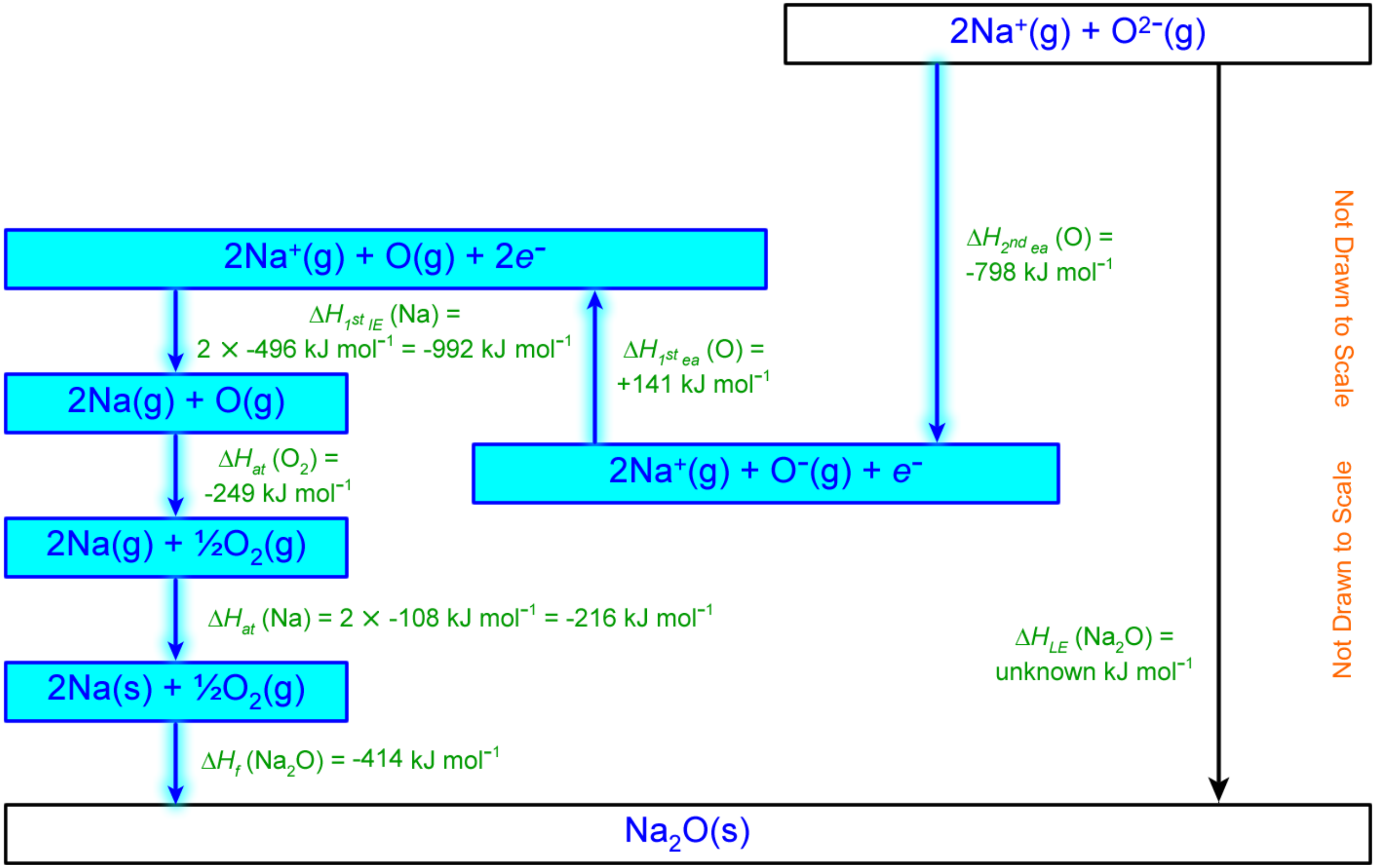
Not Drawn to Scale

- It is not possible to go directly from $2Na^+(g) + O^{2-}(g)$ to $Na_2O(s)$.

Advance Concepts – Born-Haber Cycles

Born-Haber Cycle to Calculate ΔH_{LE} for Na_2O



$2Na^+(g) + O^{2-}(g)$

$2Na^+(g) + O(g) + 2e^-$

$\Delta H_{1st\ IE}$ (Na) = 2×-496 kJ mol^{-1} = -992 kJ mol^{-1}

$2Na(g) + O(g)$

ΔH_{at} (O_2) = -249 kJ mol^{-1}

$2Na(g) + \frac{1}{2}O_2(g)$

ΔH_{at} (Na) = 2×-108 kJ mol^{-1} = -216 kJ mol^{-1}

$2Na(s) + \frac{1}{2}O_2(g)$

ΔH_f (Na_2O) = -414 kJ mol^{-1}

$Na_2O(s)$

$\Delta H_{1st\ ea}$ (O) = $+141$ kJ mol^{-1}

$2Na^+(g) + O^-(g) + e^-$

$\Delta H_{2nd\ ea}$ (O) = -798 kJ mol^{-1}

ΔH_{LE} (Na_2O) = unknown kJ mol^{-1}

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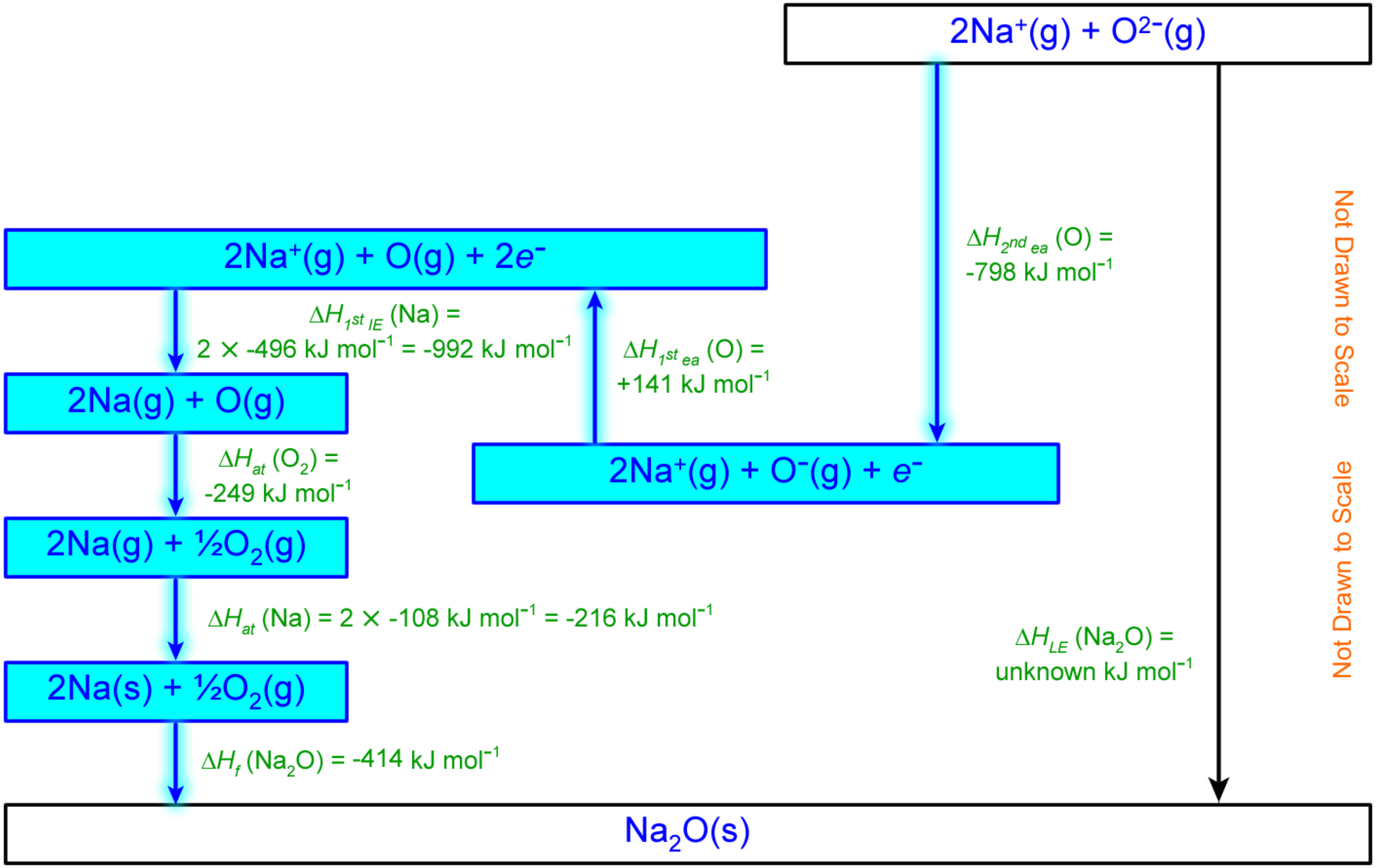
Not Drawn to Scale

- Follow the direction of the blue arrows through the cycle.
- Some arrows have to be reversed. Reversing an arrow changes the sign (+ to – and – to +) of the enthalpy change that is associated with it.

Advance Concepts – Born-Haber Cycles

Born-Haber Cycle to Calculate ΔH_{LE} for Na_2O



$2Na^+(g) + O^{2-}(g)$

$\Delta H_{2nd\ ea}$ (O) = -798 kJ mol^{-1}

$2Na^+(g) + O(g) + 2e^-$

$\Delta H_{1st\ IE}$ (Na) = 2 × -496 kJ mol^{-1} = -992 kJ mol^{-1}

$\Delta H_{1st\ ea}$ (O) = +141 kJ mol^{-1}

$2Na(g) + O(g)$

$2Na^+(g) + O^-(g) + e^-$

ΔH_{at} (O_2) = -249 kJ mol^{-1}

$2Na(g) + \frac{1}{2}O_2(g)$

ΔH_{at} (Na) = 2 × -108 kJ mol^{-1} = -216 kJ mol^{-1}

$2Na(s) + \frac{1}{2}O_2(g)$

ΔH_f (Na_2O) = -414 kJ mol^{-1}

ΔH_{LE} (Na_2O) = unknown kJ mol^{-1}

$Na_2O(s)$

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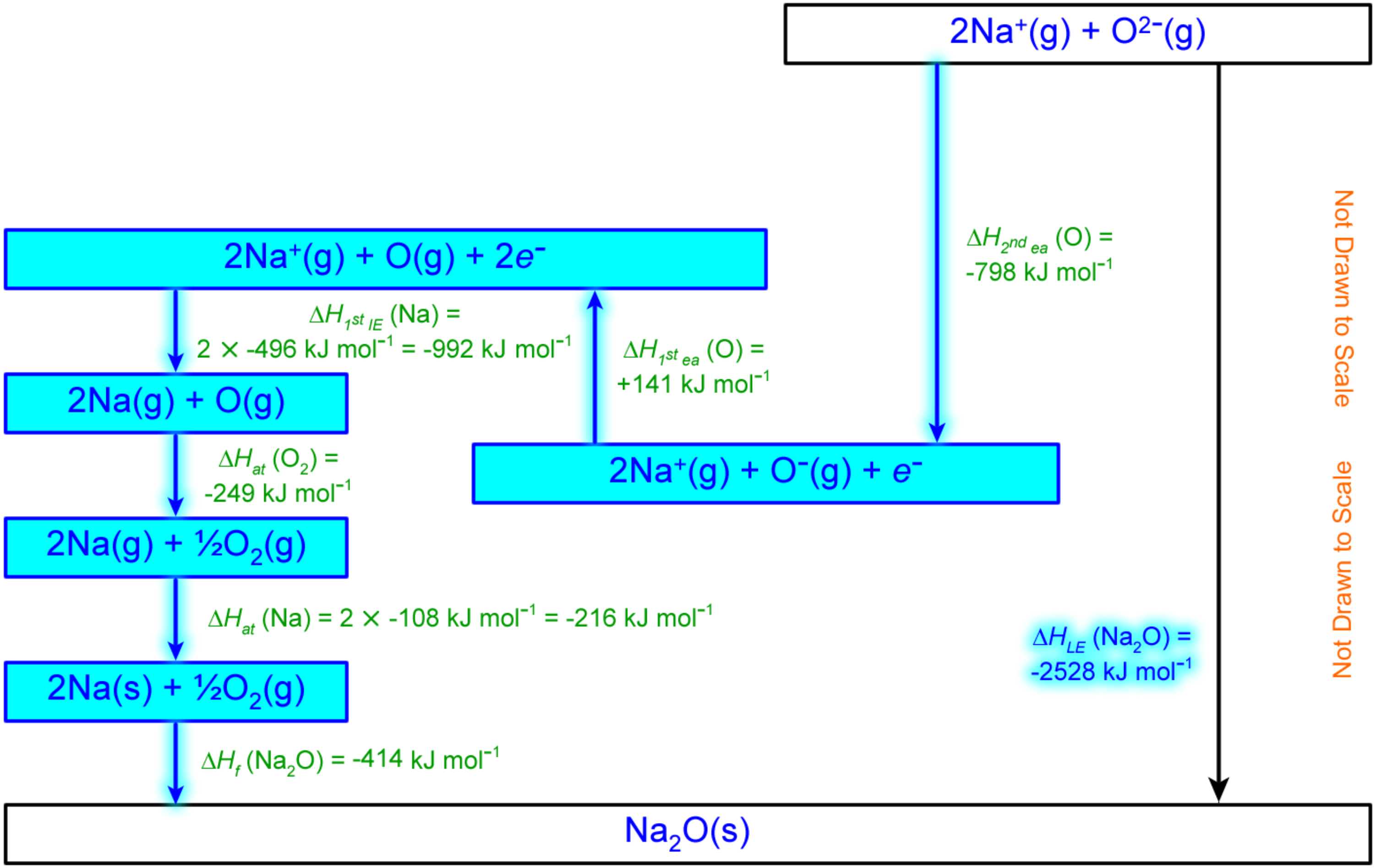
Not Drawn to Scale

$$\Delta H_{LE}\ (Na_2O) = (-798) + (+141) + (-992) + (-249) + (-216) + (-414)$$

Advance Concepts – Born-Haber Cycles

Born-Haber Cycle to Calculate ΔH_{LE} for Na_2O



$2Na^+(g) + O^{2-}(g)$

$\Delta H_{2nd\ ea}$ (O) = -798 kJ mol^{-1}

$2Na^+(g) + O(g) + 2e^-$

$\Delta H_{1st\ IE}$ (Na) = 2 × -496 kJ mol^{-1} = -992 kJ mol^{-1}

$\Delta H_{1st\ ea}$ (O) = +141 kJ mol^{-1}

$2Na(g) + O(g)$

$2Na^+(g) + O^-(g) + e^-$

ΔH_{at} (O_2) = -249 kJ mol^{-1}

$2Na(g) + ½O_2(g)$

ΔH_{at} (Na) = 2 × -108 kJ mol^{-1} = -216 kJ mol^{-1}

ΔH_{LE} (Na_2O) = -2528 kJ mol^{-1}

$2Na(s) + ½O_2(g)$

ΔH_f (Na_2O) = -414 kJ mol^{-1}

$Na_2O(s)$

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$$\Delta H_{LE}\ (Na_2O) = (-798) + (+141) + (-992) + (-249) + (-216) + (-414)$$

$$\Delta H_{LE}\ (Na_2O) = -2528\ \text{kJ mol}^{-1}$$

Advance Concepts – Born-Haber Cycles

- Use the data given below to calculate the *second ionization energy of magnesium*.

$Mg(s) + {}^{1}/_{2}O_2(g) \rightarrow MgO(s) \quad \Delta H_{formation} = -602 \text{ kJ mol}^{-1}$

$Mg(s) \rightarrow Mg(g) \quad \Delta H_{atomisation} = +148 \text{ kJ mol}^{-1}$

${}^{1}/_{2}O_2(g) \rightarrow O(g) \quad \Delta H_{atomisation} = +249 \text{ kJ mol}^{-1}$

$Mg(g) \rightarrow Mg^{+}(g) + e^{-} \quad \Delta H_{first\ ionisation} = +738 \text{ kJ mol}^{-1}$

$Mg^{+}(g) \rightarrow Mg^{2+}(g) + e^{-} \quad \Delta H_{second\ ionisation} = \text{unknown kJ mol}^{-1}$

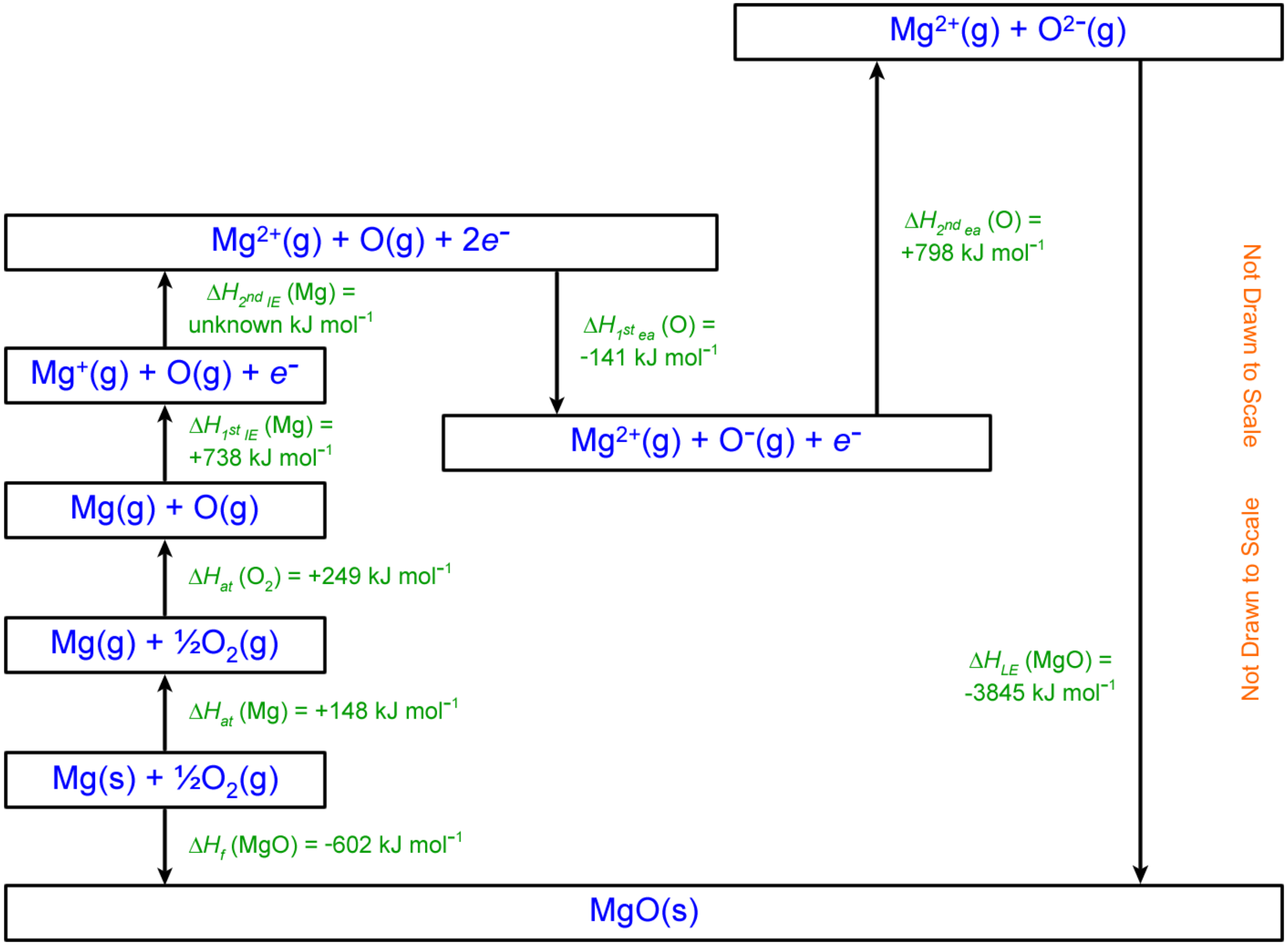
$O(g) + e^{-} \rightarrow O^{-}(g) \quad \Delta H_{first\ electron\ affinity} = -141 \text{ kJ mol}^{-1}$

$O^{-}(g) + e^{-} \rightarrow O^{2-}(g) \quad \Delta H_{second\ electron\ affinity} = +798 \text{ kJ mol}^{-1}$

$Mg^{2+}(g) + O^{2-}(g) \rightarrow MgO(s) \quad \Delta H_{lattice\ enthalpy} = -3845 \text{ kJ mol}^{-1}$

Advance Concepts – Born-Haber Cycles

Born-Haber Cycle to Calculate $\Delta H_{2nd\ IE}$ for Mg



$Mg^{2+}(g) + O^{2-}(g)$

$Mg^{2+}(g) + O(g) + 2e^-$

$\Delta H_{2nd\ ea}$ (O) = +798 kJ mol^{-1}

$\Delta H_{2nd\ IE}$ (Mg) = unknown kJ mol^{-1}

$Mg^+(g) + O(g) + e^-$

$\Delta H_{1st\ ea}$ (O) = -141 kJ mol^{-1}

$\Delta H_{1st\ IE}$ (Mg) = +738 kJ mol^{-1}

$Mg^{2+}(g) + O^-(g) + e^-$

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$Mg(g) + O(g)$

ΔH_{at} (O_2) = +249 kJ mol^{-1}

$Mg(g) + \frac{1}{2}O_2(g)$

ΔH_{LE} (MgO) = -3845 kJ mol^{-1}

ΔH_{at} (Mg) = +148 kJ mol^{-1}

$Mg(s) + \frac{1}{2}O_2(g)$

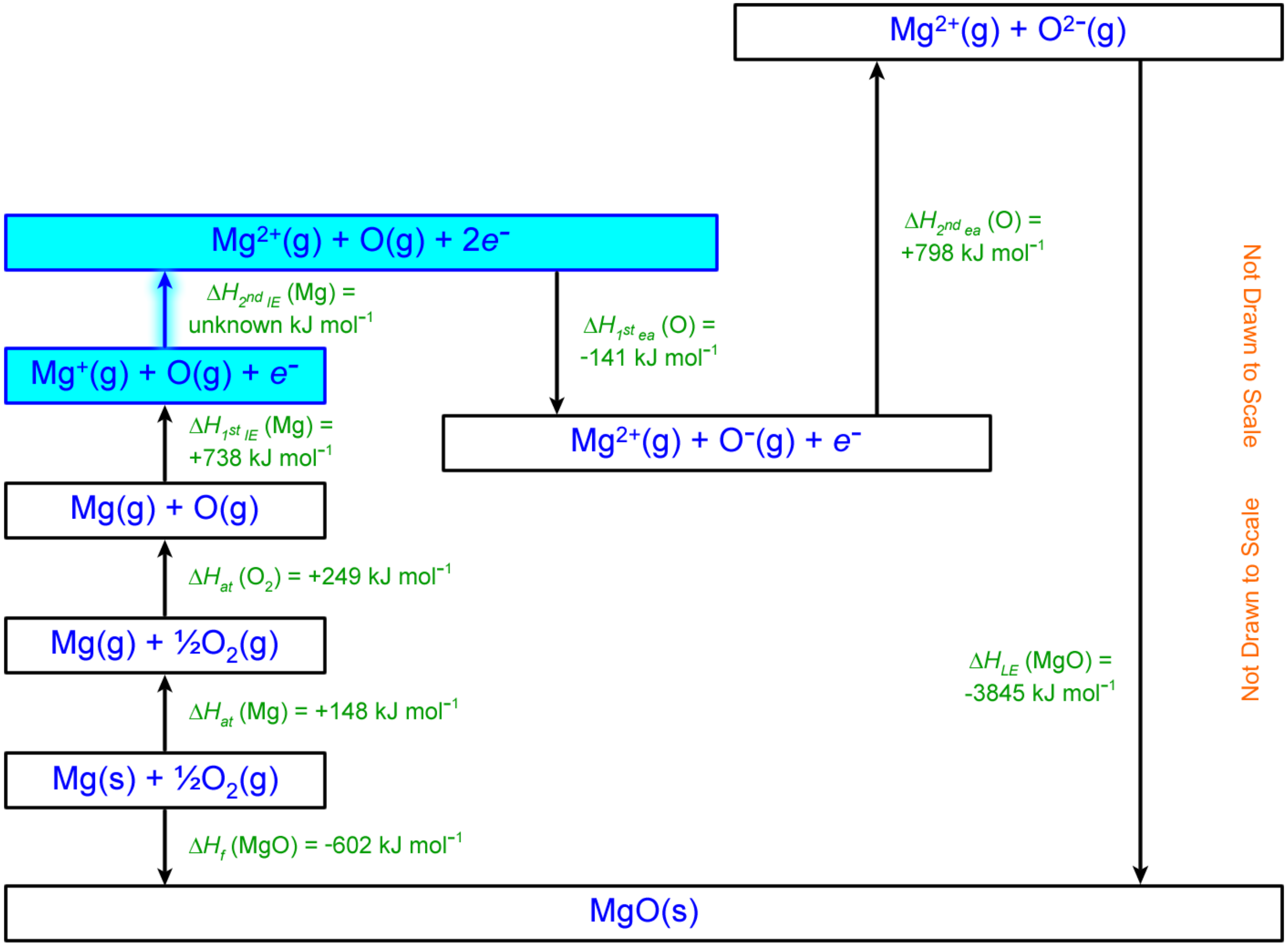
ΔH_f (MgO) = -602 kJ mol^{-1}

$MgO(s)$

- Calculate the second ionization energy for Mg(s).

Advance Concepts – Born-Haber Cycles

Born-Haber Cycle to Calculate $\Delta H_{2nd\ IE}$ for Mg



$Mg^{2+}(g) + O^{2-}(g)$

$\Delta H_{2nd\ ea}$ (O) = +798 kJ mol^{-1}

$Mg^{2+}(g) + O(g) + 2e^{-}$

$\Delta H_{2nd\ IE}$ (Mg) = unknown kJ mol^{-1}

$\Delta H_{1st\ ea}$ (O) = -141 kJ mol^{-1}

$Mg^{+}(g) + O(g) + e^{-}$

$\Delta H_{1st\ IE}$ (Mg) = +738 kJ mol^{-1}

$Mg^{2+}(g) + O^{-}(g) + e^{-}$

$Mg(g) + O(g)$

ΔH_{at} (O_2) = +249 kJ mol^{-1}

$Mg(g) + ½O_2(g)$

ΔH_{LE} (MgO) = -3845 kJ mol^{-1}

ΔH_{at} (Mg) = +148 kJ mol^{-1}

$Mg(s) + ½O_2(g)$

ΔH_f (MgO) = -602 kJ mol^{-1}

MgO(s)

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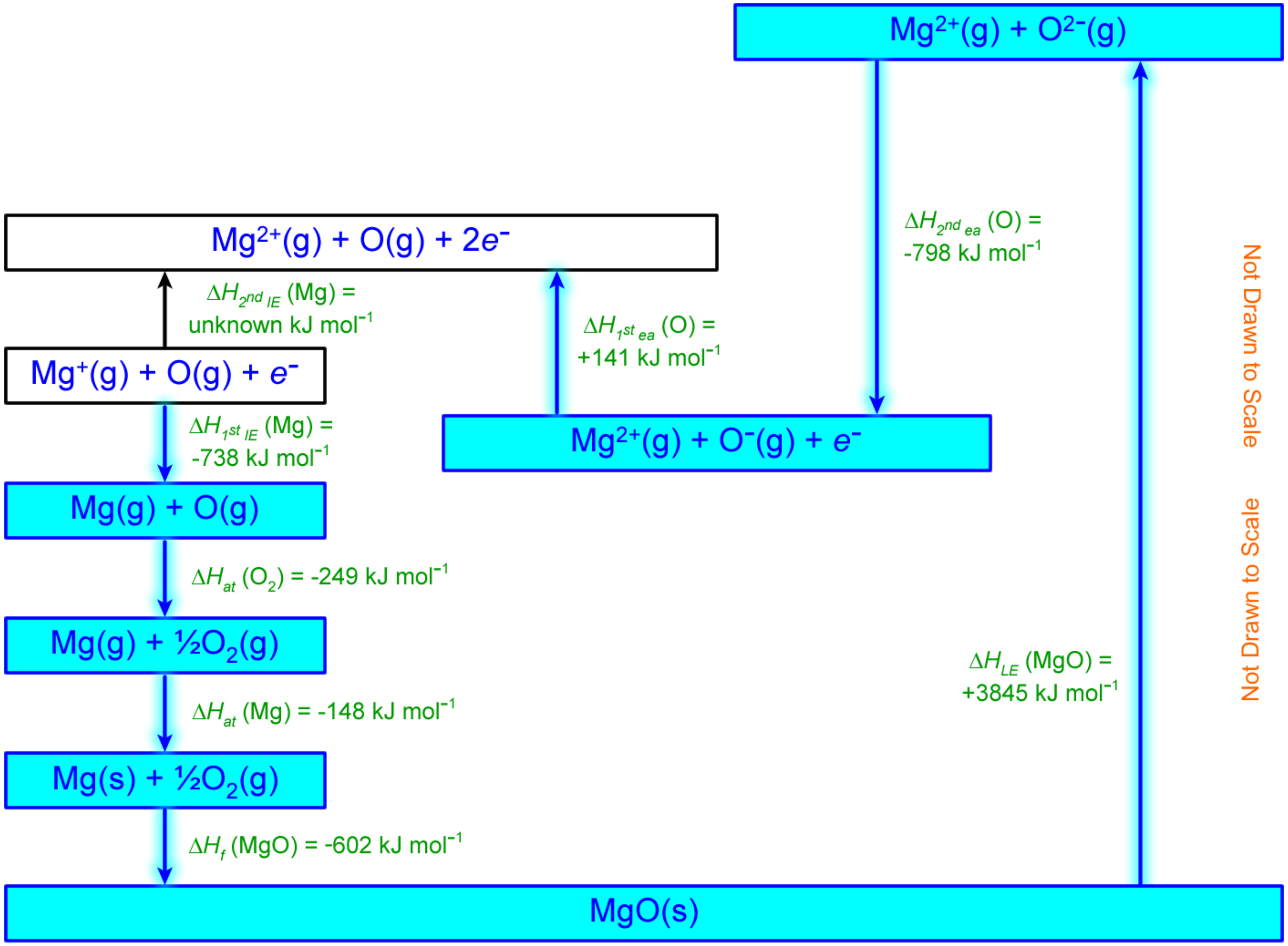
Not Drawn to Scale

- It is not possible to go directly from $Mg^{+}(g)$ to $Mg^{2+}(g)$.

Advance Concepts – Born-Haber Cycles

Born-Haber Cycle to Calculate $\Delta H_{2nd\ IE}$ for Mg



$Mg^{2+}(g) + O^{2-}(g)$

$Mg^{2+}(g) + O(g) + 2e^-$

$\Delta H_{2nd\ ea}$ (O) = -798 kJ mol^{-1}

$\Delta H_{2nd\ IE}$ (Mg) = unknown kJ mol^{-1}

$\Delta H_{1st\ ea}$ (O) = +141 kJ mol^{-1}

$Mg^{+}(g) + O(g) + e^-$

$Mg^{2+}(g) + O^{-}(g) + e^-$

$\Delta H_{1st\ IE}$ (Mg) = -738 kJ mol^{-1}

$Mg(g) + O(g)$

ΔH_{at} (O_2) = -249 kJ mol^{-1}

$Mg(g) + ½O_2(g)$

ΔH_{at} (Mg) = -148 kJ mol^{-1}

$Mg(s) + ½O_2(g)$

ΔH_f (MgO) = -602 kJ mol^{-1}

ΔH_{LE} (MgO) = +3845 kJ mol^{-1}

Not Drawn to Scale

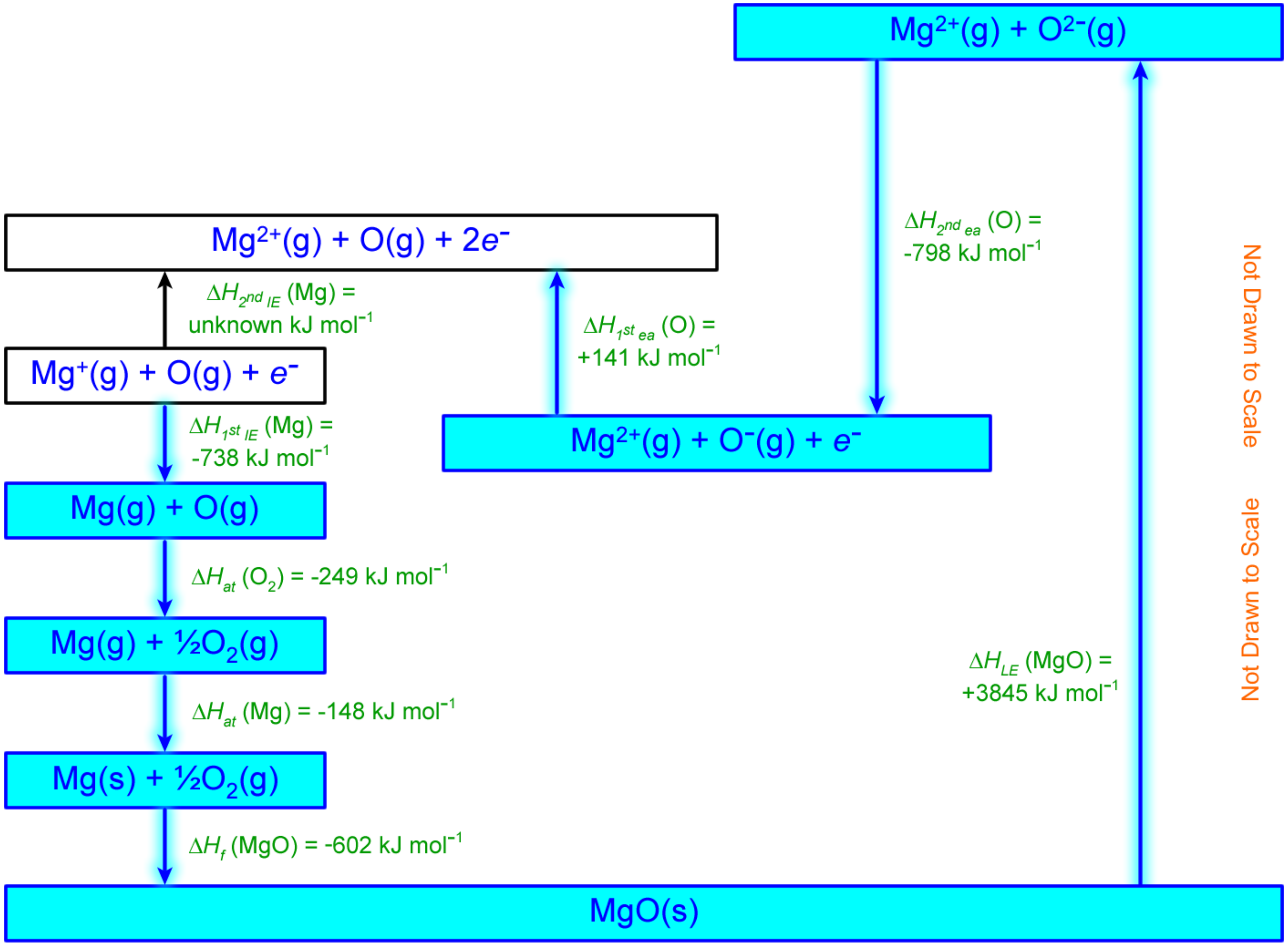
Not Drawn to Scale

MgO(s)

- Follow the direction of the blue arrows through the cycle.
- Some arrows have to be reversed. Reversing an arrow changes the sign (+ to – and – to +) of the enthalpy change that is associated with it.

Advance Concepts – Born-Haber Cycles

Born-Haber Cycle to Calculate $\Delta H_{2nd\ IE}$ for Mg



$Mg^{2+}(g) + O^{2-}(g)$

$\Delta H_{2nd\ ea}$ (O) = -798 kJ mol^{-1}

$Mg^{2+}(g) + O(g) + 2e^{-}$

$\Delta H_{2nd\ IE}$ (Mg) = unknown kJ mol^{-1}

$\Delta H_{1st\ ea}$ (O) = +141 kJ mol^{-1}

$Mg^{+}(g) + O(g) + e^{-}$

$\Delta H_{1st\ IE}$ (Mg) = -738 kJ mol^{-1}

$Mg^{2+}(g) + O^{-}(g) + e^{-}$

$Mg(g) + O(g)$

ΔH_{at} (O_2) = -249 kJ mol^{-1}

$Mg(g) + \frac{1}{2}O_2(g)$

ΔH_{at} (Mg) = -148 kJ mol^{-1}

$Mg(s) + \frac{1}{2}O_2(g)$

ΔH_f (MgO) = -602 kJ mol^{-1}

ΔH_{LE} (MgO) = +3845 kJ mol^{-1}

Not Drawn to Scale

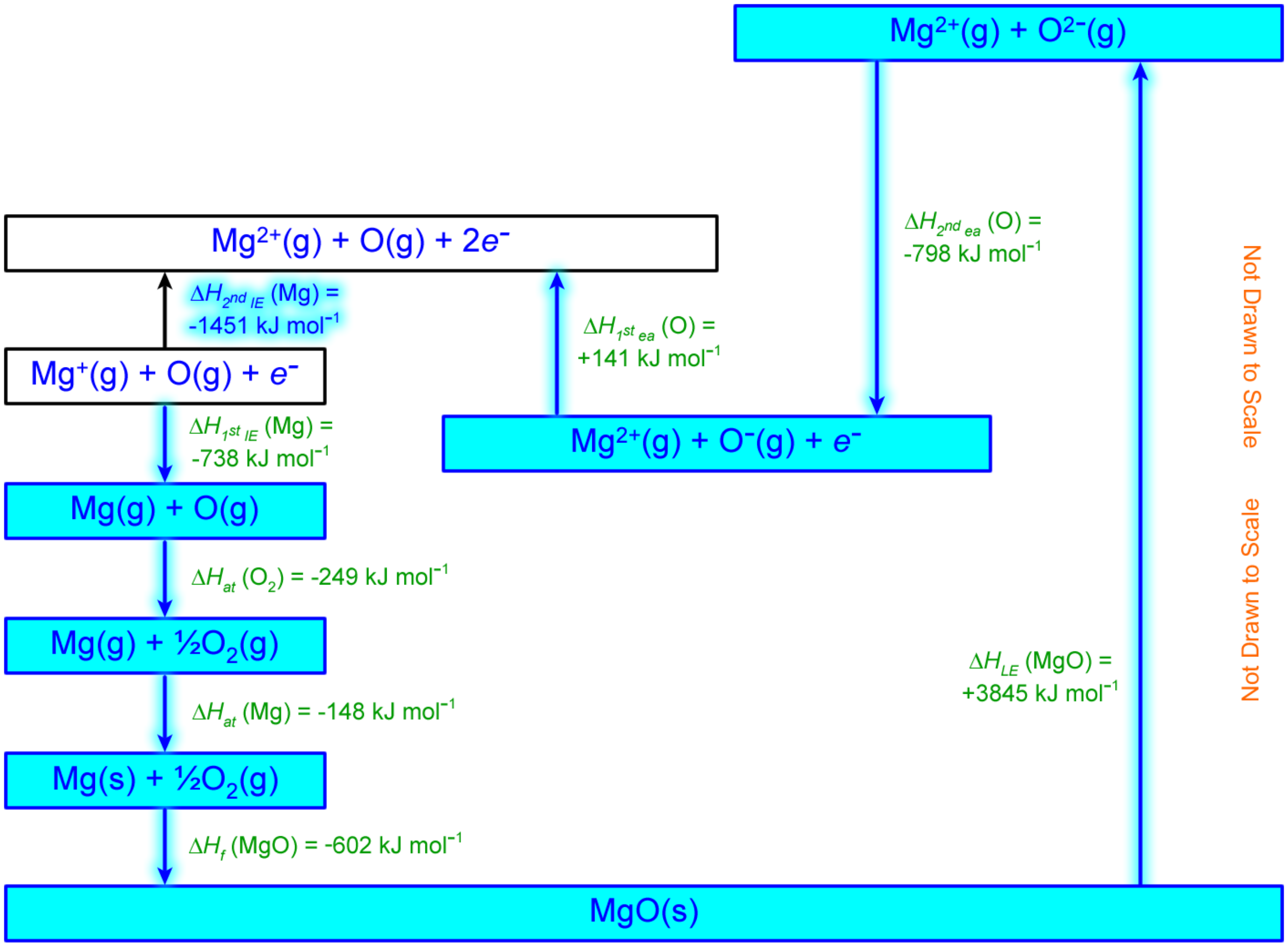
Not Drawn to Scale

$MgO(s)$

$$\Delta H_{2nd\ IE}\ (Mg) = (-738) + (-249) + (-148) + (-602) + (+3845) + (-798) + (+141)$$

Advance Concepts – Born-Haber Cycles

Born-Haber Cycle to Calculate $\Delta H_{2nd\ IE}$ for Mg



$Mg^{2+}(g) + O^{2-}(g)$

$\Delta H_{2nd\ ea}$ (O) = -798 kJ mol^{-1}

$Mg^{2+}(g) + O(g) + 2e^{-}$

$\Delta H_{2nd\ IE}$ (Mg) = -1451 kJ mol^{-1}

$\Delta H_{1st\ ea}$ (O) = +141 kJ mol^{-1}

$Mg^{+}(g) + O(g) + e^{-}$

$Mg^{2+}(g) + O^{-}(g) + e^{-}$

$\Delta H_{1st\ IE}$ (Mg) = -738 kJ mol^{-1}

$Mg(g) + O(g)$

ΔH_{at} (O_2) = -249 kJ mol^{-1}

$Mg(g) + ½O_2(g)$

ΔH_{LE} (MgO) = +3845 kJ mol^{-1}

ΔH_{at} (Mg) = -148 kJ mol^{-1}

$Mg(s) + ½O_2(g)$

ΔH_f (MgO) = -602 kJ mol^{-1}

MgO(s)

Not Drawn to Scale

Not Drawn to Scale

$$\Delta H_{2nd\ IE} \text{ (Mg)} = (-738) + (-249) + (-148) + (-602) + (+3845) + (-798) + (+141)$$

$$\Delta H_{2nd\ IE} \text{ (Mg)} = +1451 \text{ kJ mol}^{-1}$$

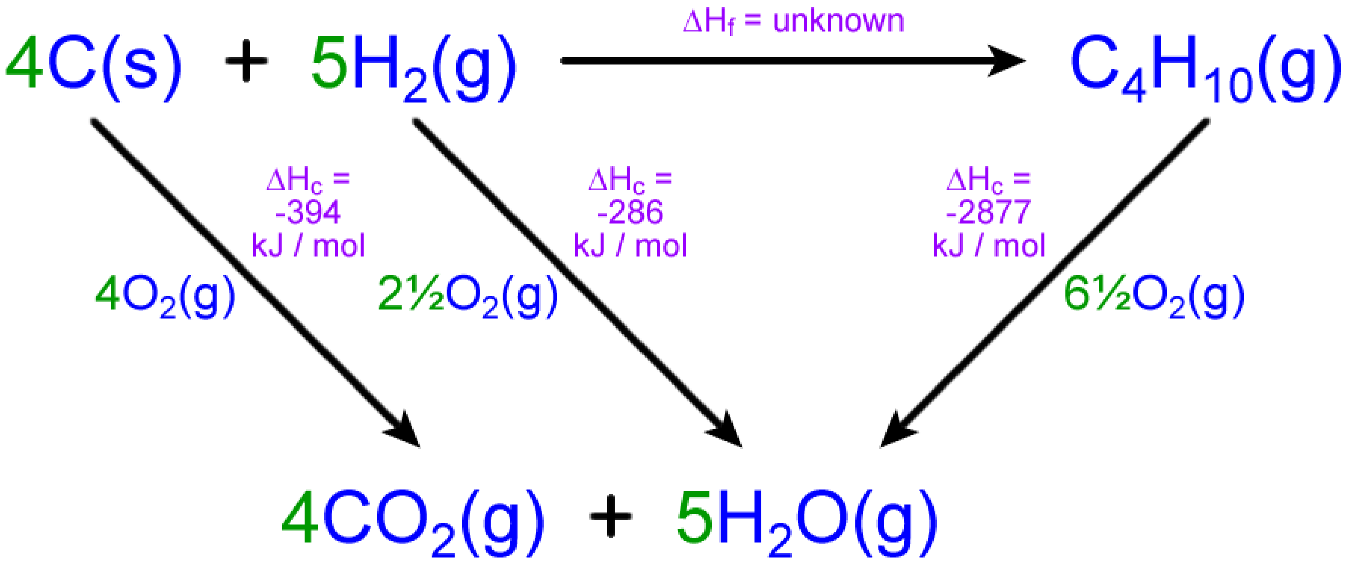
Advance Concepts – Hess Cycles

What is *Hess's Law*?

Why are *Hess Cycles* useful?

Advance Concepts – Hess Cycles

- Hess's Law states that the *total enthalpy change* during the *complete course* of a chemical reaction is the *same*, whether the reaction takes place in *one step*, or it takes place in *several steps*.



$4C(s) + 5H_2(g)$

ΔH_f = unknown

$C_4H_{10}(g)$

$4O_2(g)$

ΔH_c = -394 kJ / mol

$2\frac{1}{2}O_2(g)$

ΔH_c = -286 kJ / mol

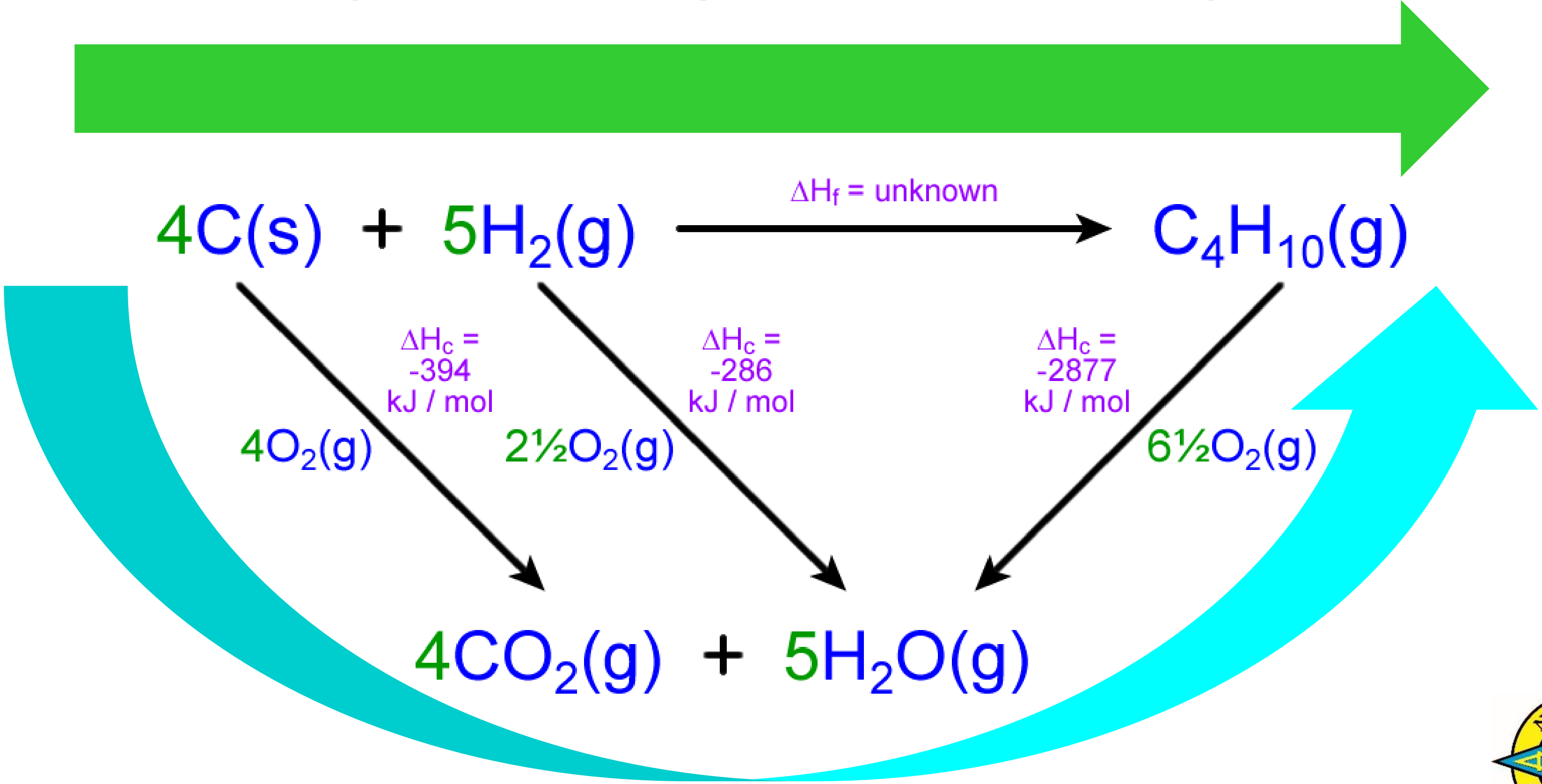
ΔH_c = -2877 kJ / mol

$6\frac{1}{2}O_2(g)$

$4CO_2(g) + 5H_2O(g)$

Advance Concepts – Hess Cycles

- Hess's Law states that the *total enthalpy change* during the *complete course* of a chemical reaction is the *same*, whether the reaction takes place in *one step*, or it takes place in *several steps*.



$$4C(s) + 5H_2(g) \xrightarrow{\Delta H_f = \text{unknown}} C_4H_{10}(g)$$

ΔH_c = -394 kJ / mol

$4O_2(g)$

ΔH_c = -286 kJ / mol

$2\tfrac{1}{2}O_2(g)$

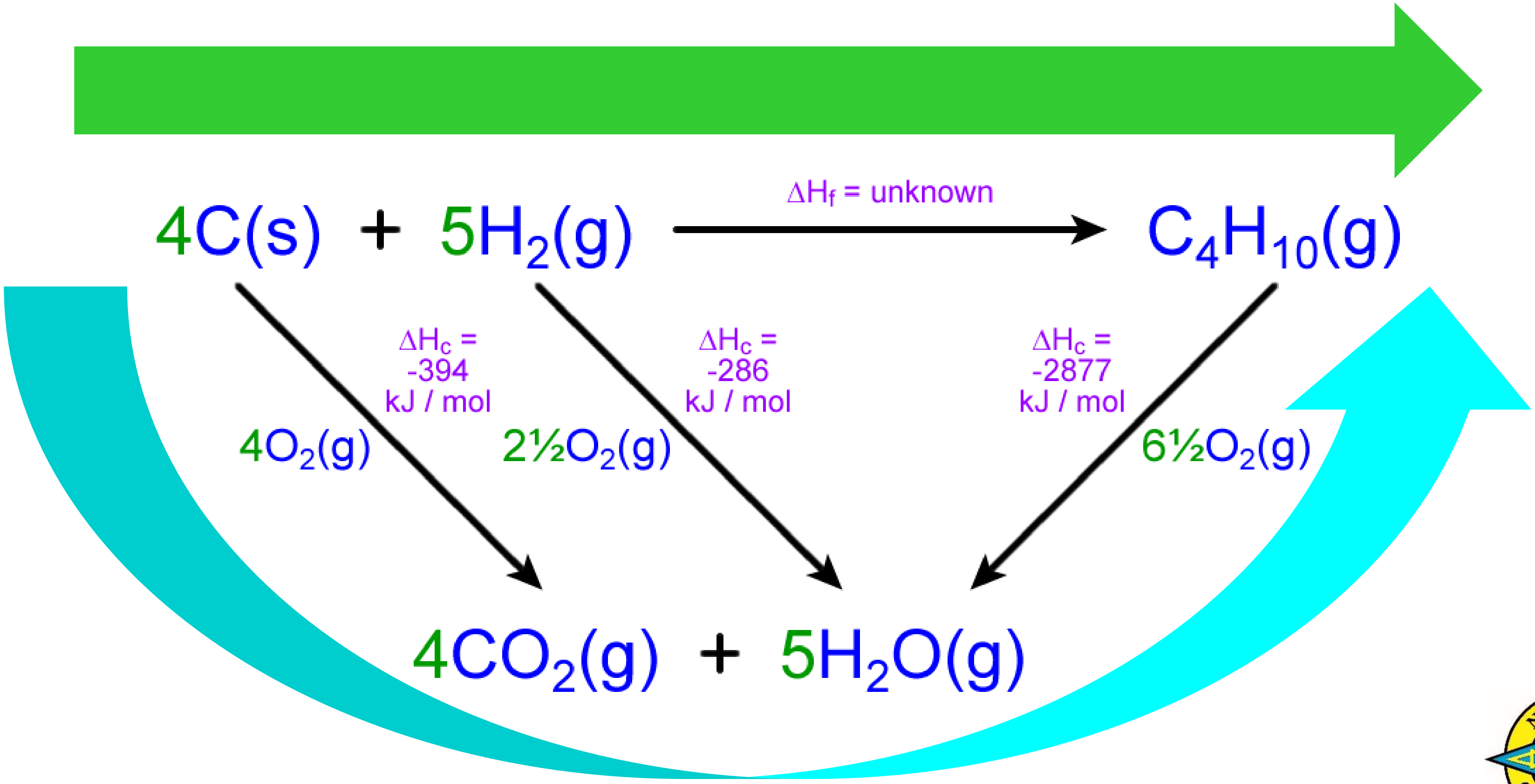
ΔH_c = -2877 kJ / mol

$6\tfrac{1}{2}O_2(g)$

$$4CO_2(g) + 5H_2O(g)$$

Advance Concepts – Hess Cycles

- Hess's Law maybe rephrased to say that the enthalpy change of a chemical reaction is *independent* of the path taken from the reactants to the products.



$4C(s) + 5H_2(g) \xrightarrow{\Delta H_f = \text{unknown}} C_4H_{10}(g)$

$4O_2(g)$

$\Delta H_c =$ -394 kJ / mol

$2\frac{1}{2}O_2(g)$

$\Delta H_c =$ -286 kJ / mol

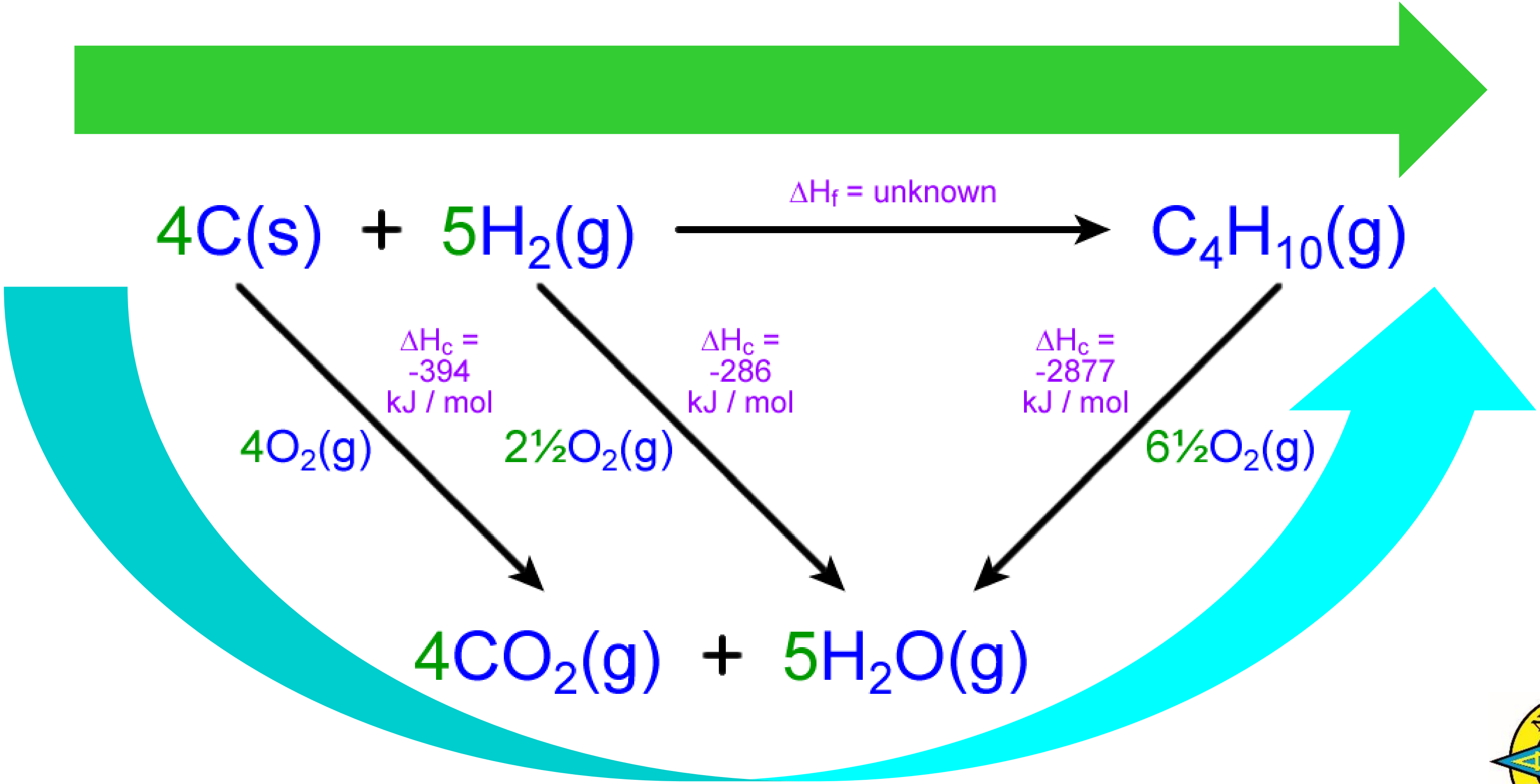
$6\frac{1}{2}O_2(g)$

$\Delta H_c =$ -2877 kJ / mol

$4CO_2(g) + 5H_2O(g)$

Advance Concepts – Hess Cycles

- Hess's Law is based upon the principle that *energy must be conserved* during a chemical reaction (energy can not be created or destroyed, it can only be converted from one form to another).



$$4C(s) + 5H_2(g) \xrightarrow{\Delta H_f = \text{unknown}} C_4H_{10}(g)$$

ΔH_c = -394 kJ / mol

$4O_2(g)$

ΔH_c = -286 kJ / mol

$2\frac{1}{2}O_2(g)$

ΔH_c = -2877 kJ / mol

$6\frac{1}{2}O_2(g)$

$$4CO_2(g) + 5H_2O(g)$$

Advance Concepts – Hess Cycles

- Hess's Law allows the enthalpy change (ΔH) of a chemical reaction to be calculated, even when it cannot be measured directly by experiment.
- This is achieved by performing simple calculations for related reactions, the enthalpy changes for which are known.

Advance Concepts – Hess Cycles

- The enthalpy change of formation, ΔH_f, for butane, C_4H_{10}, is the enthalpy change when 1 mole of butane is formed from carbon (as graphite) and hydrogen in their standard states at room temperature and pressure.
- The enthalpy change of formation for butane cannot be measured directly in the laboratory, but it can be calculated using a Hess Cycle.
- Use the following enthalpy changes of combustion, ΔH_c, to construct an appropriate Hess Cycle, and hence calculate the enthalpy change of formation for butane:

ΔH_c for graphite = -394 kJ mol^{-1}

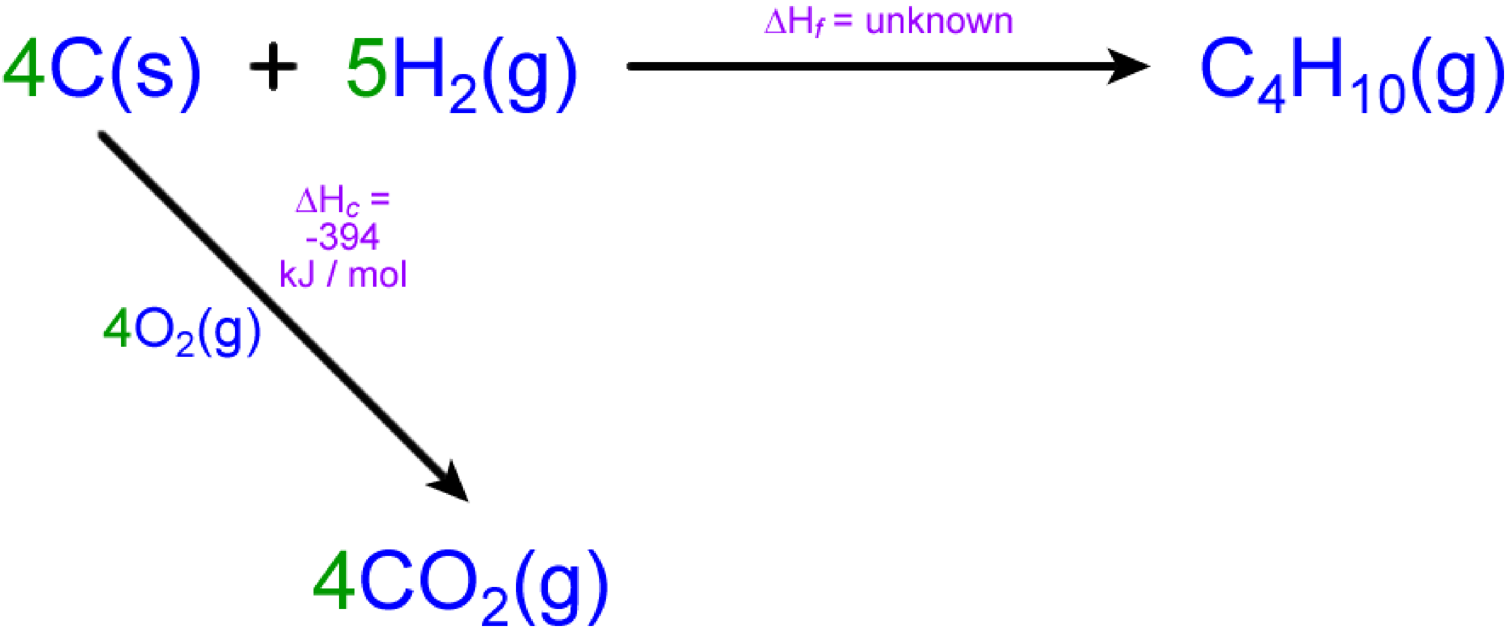
ΔH_c for hydrogen = -286 kJ mol^{-1}

ΔH_c for butane = -2877 kJ mol^{-1}

Advance Concepts – Hess Cycles

$$4C(s) + 5H_2(g) \xrightarrow{\Delta H_f = \text{unknown}} C_4H_{10}(g)$$

Advance Concepts – Hess Cycles



$$4C(s) + 5H_2(g) \xrightarrow{\Delta H_f = \text{unknown}} C_4H_{10}(g)$$

$$4C(s) \xrightarrow[4O_2(g)]{\Delta H_c = -394\ \text{kJ / mol}} 4CO_2(g)$$

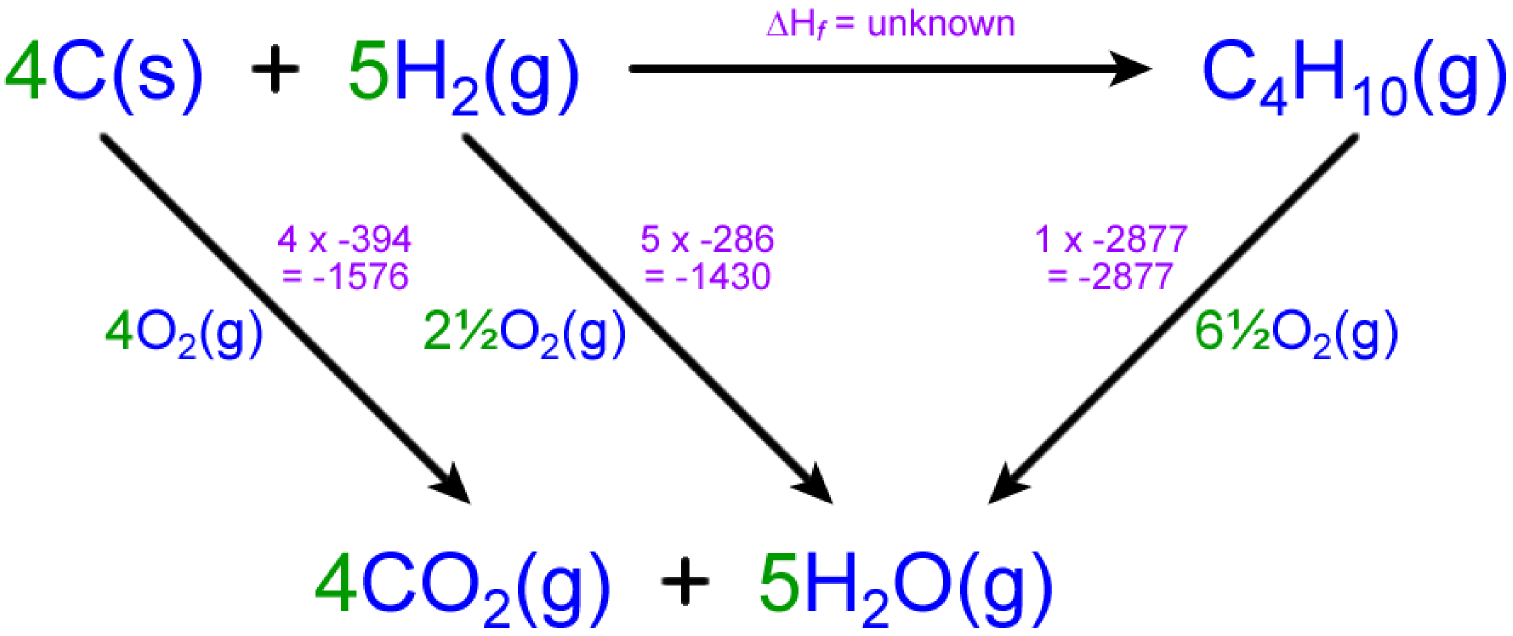
Advance Concepts – Hess Cycles

$$\begin{tikzcd}
4C(s) + \arrow[dr, "4O_2(g)"', "\Delta H_c = -394\ \text{kJ / mol}"] & 5H_2(g) \arrow[rr, "\Delta H_f = \text{unknown}"] \arrow[dr, "2\frac{1}{2}O_2(g)"', "\Delta H_c = -286\ \text{kJ / mol}"] & & C_4H_{10}(g) \\
& 4CO_2(g) + & 5H_2O(g) &
\end{tikzcd}$$

Advance Concepts – Hess Cycles

$$\begin{tikzcd}
4C(s) \arrow[dr, "\Delta H_c = -394 \text{ kJ / mol}", "4O_2(g)"'] & + & 5H_2(g) \arrow[rr, "\Delta H_f = \text{unknown}"] \arrow[dr, "\Delta H_c = -286 \text{ kJ / mol}", "2\tfrac{1}{2}O_2(g)"'] & & C_4H_{10}(g) \arrow[dl, "\Delta H_c = -2877 \text{ kJ / mol}"', "6\tfrac{1}{2}O_2(g)"] \\
& 4CO_2(g) & + & 5H_2O(g) &
\end{tikzcd}$$

Advance Concepts – Hess Cycles



$4C(s) + 5H_2(g)$

ΔH_f = unknown

$C_4H_{10}(g)$

$4O_2(g)$

4 x -394
= -1576

$2\frac{1}{2}O_2(g)$

5 x -286
= -1430

1 x -2877
= -2877

$6\frac{1}{2}O_2(g)$

$4CO_2(g) + 5H_2O(g)$

Advance Concepts – Hess Cycles

- Reverse the direction of the arrow between C_4H_{10} and CO_2 / H_2O to ensure that the Hess Cycle moves in the correct direction, from reactants to products:

$4C(s) + 5H_2(g) \rightarrow 4CO_2(g) + 5H_2O(g) \rightarrow C_4H_{10}(g)$

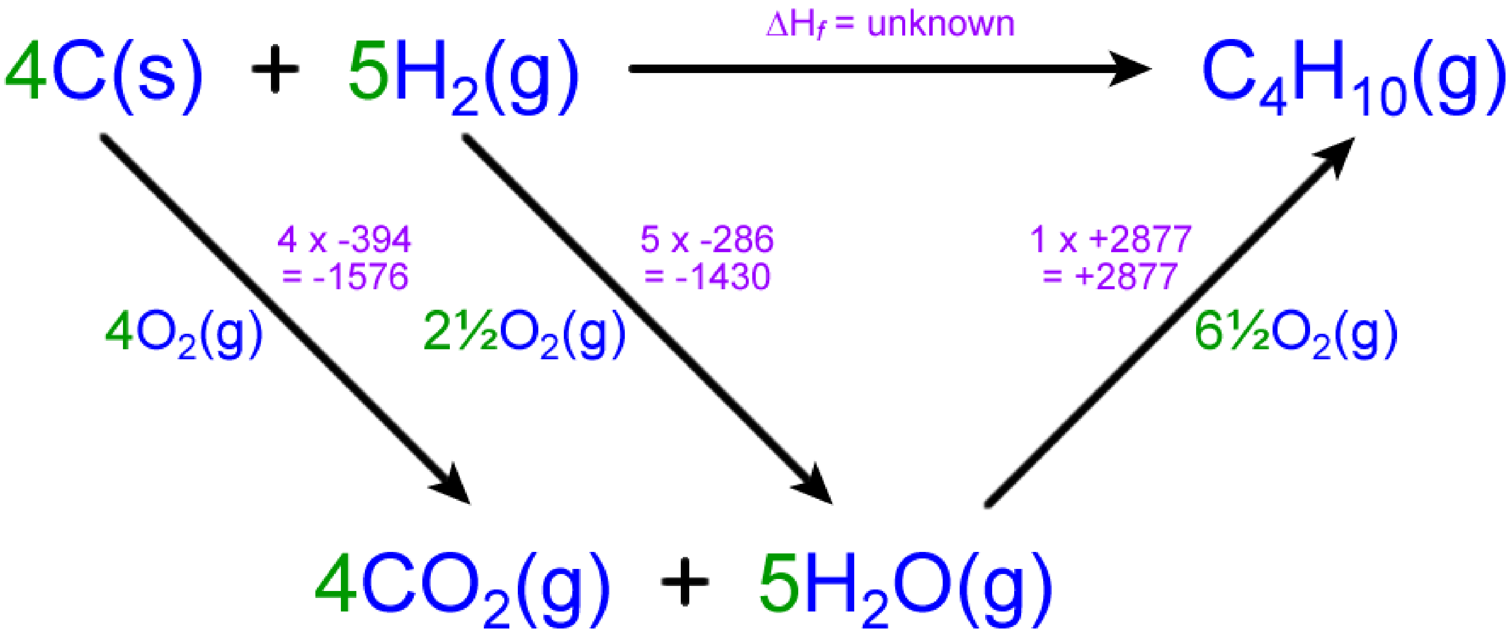
Note: If the direction of an arrow is reversed, then the sign of the enthalpy change associated with the arrow is also changed (what was once an exothermic reaction is now an endothermic reaction).

- The enthalpy change of formation for butane is then the sum of the enthalpy changes leading from $4C(s) + 5H_2(g)$ to $C_4H_{10}(g)$.

Advance Concepts – Hess Cycles

$$\begin{tikzcd}
4C(s) \arrow[dr, "4 \times -394 = -1576", "4O_2(g)"'] & + & 5H_2(g) \arrow[rrr, "\Delta H_f = \text{unknown}"] \arrow[dr, "5 \times -286 = -1430", "2\tfrac{1}{2}O_2(g)"'] & & & C_4H_{10}(g) \arrow[dll, "1 \times -2877 = -2877"', "6\tfrac{1}{2}O_2(g)"] \\
& 4CO_2(g) & + & 5H_2O(g) & &
\end{tikzcd}$$

Advance Concepts – Hess Cycles



$4C(s) + 5H_2(g)$

ΔH_f = unknown

$C_4H_{10}(g)$

$4O_2(g)$

4 x -394
= -1576

$2\frac{1}{2}O_2(g)$

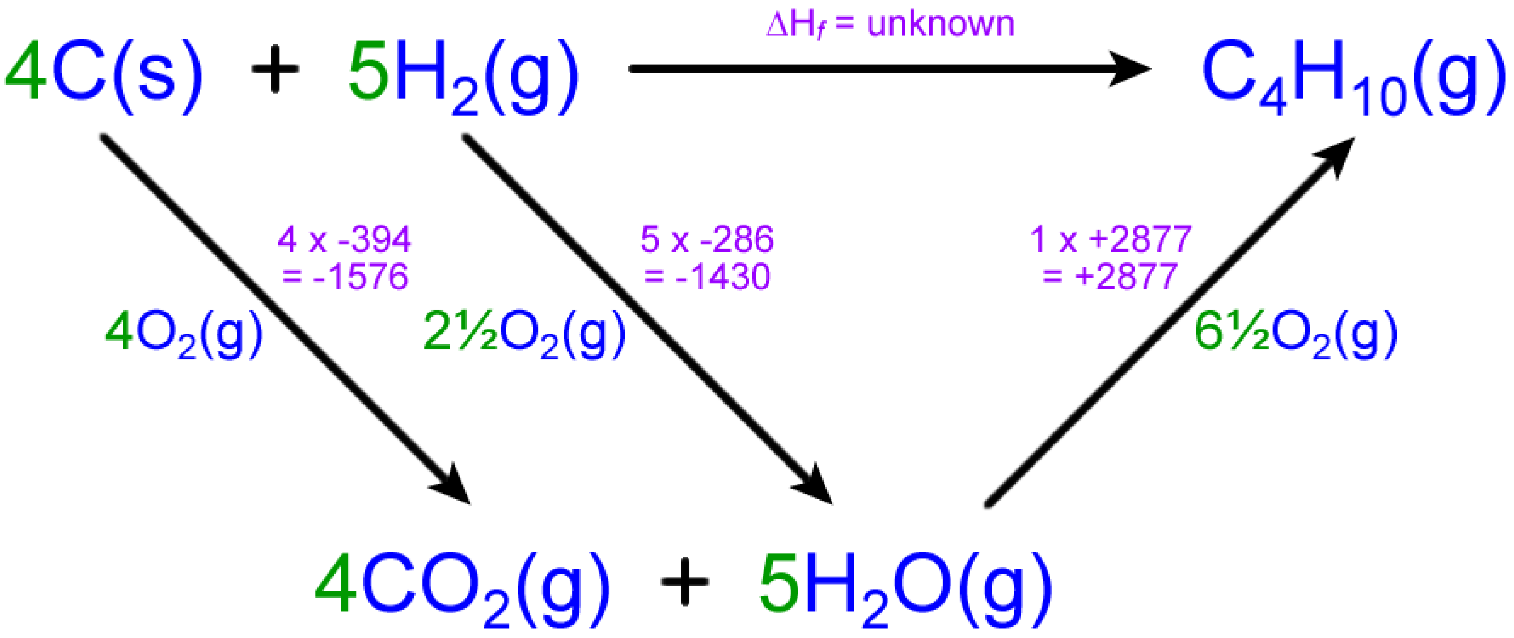
5 x -286
= -1430

1 x +2877
= +2877

$6\frac{1}{2}O_2(g)$

$4CO_2(g) + 5H_2O(g)$

Advance Concepts – Hess Cycles



ΔH_f = unknown

$4C(s) + 5H_2(g) \longrightarrow C_4H_{10}(g)$

4 x -394
= -1576

$4O_2(g)$

5 x -286
= -1430

$2\frac{1}{2}O_2(g)$

1 x +2877
= +2877

$6\frac{1}{2}O_2(g)$

$4CO_2(g) + 5H_2O(g)$

$\Delta H_f\ C_4H_{10} = (-1576) + (-1430) + (+2877)$

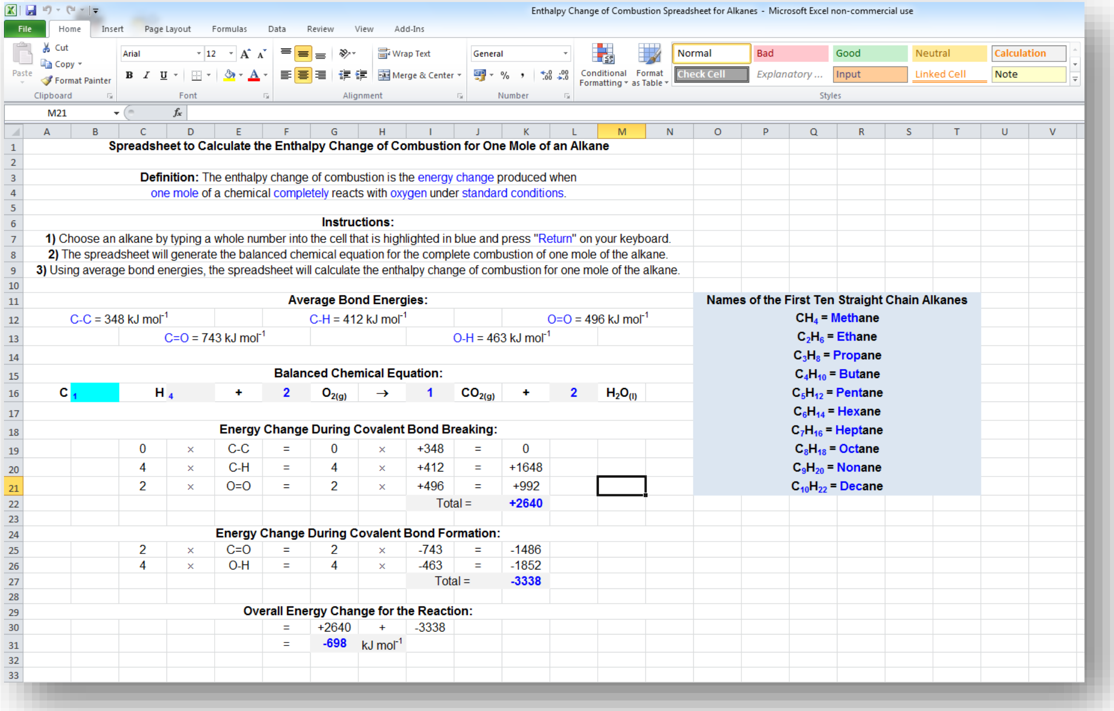
$= -129\ \text{kJ mol}^{-1}$

Alkane Combustion Spreadsheet

Can enthalpy change calculations be automated 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:

$C_1 H_4 + 2\ O_{2(g)} \rightarrow 1\ CO_{2(g)} + 2\ H_2O_{(l)}$

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**

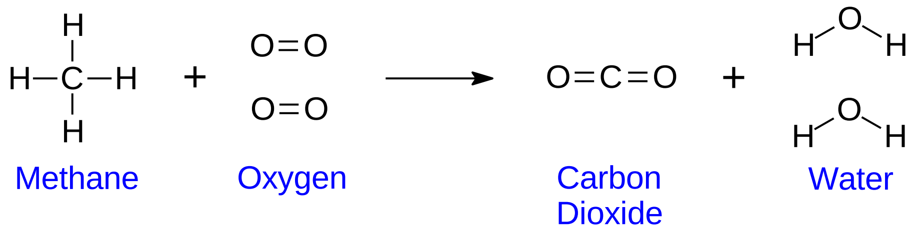
Overall Energy Change for the Reaction:

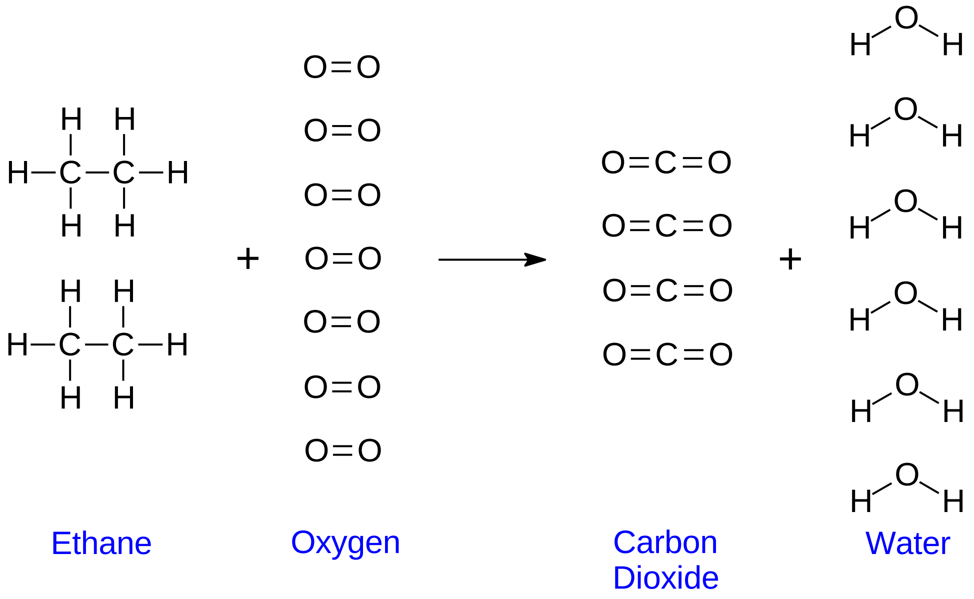
= +2640 + -3338
= **-698** kJ mol^{-1}

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

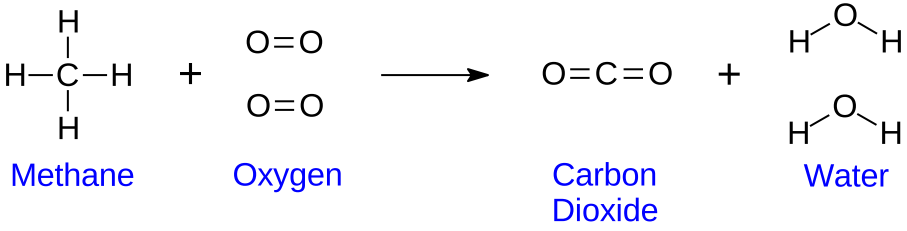
Alkane Combustion Spreadsheet

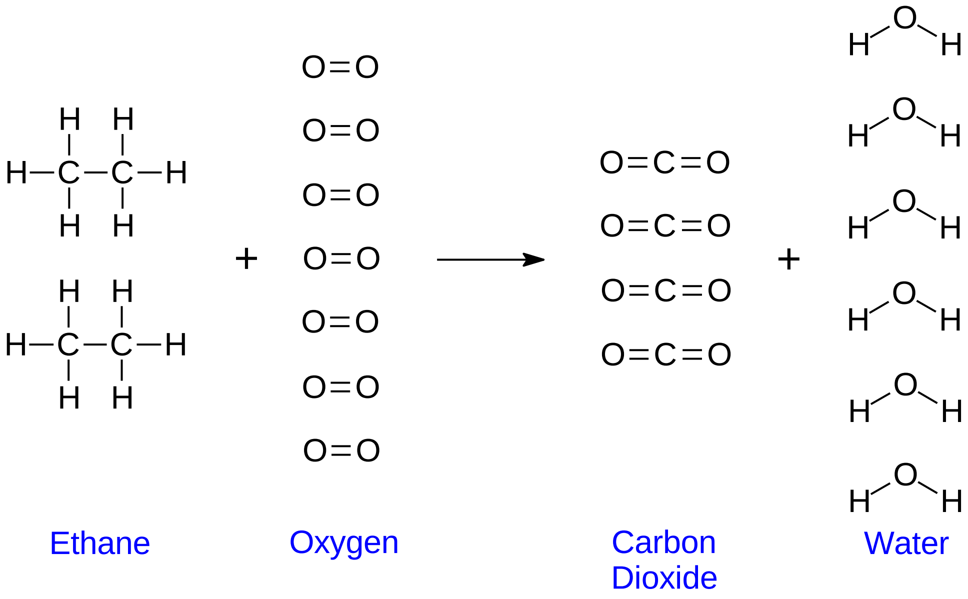




- Consider the combustion of methane and the combustion of ethane shown on the left.
- Let the number of carbon atoms = n.
- In relation to n, how many C–C bonds break?
- In relation to n, how many C–H bonds break?
- In relation to n, how many O=O bonds break?
- In relation to n, how many C=O bonds form?
- 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^{-1}.
  - Average C–H bond energy = 411 kJ mol^{-1}.
  - Average O=O bond energy = 498 kJ mol^{-1}.
  - Average C=O bond energy = 804 kJ mol^{-1}.
  - Average O–H bond energy = 463 kJ mol^{-1}.
- 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^{-1}
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

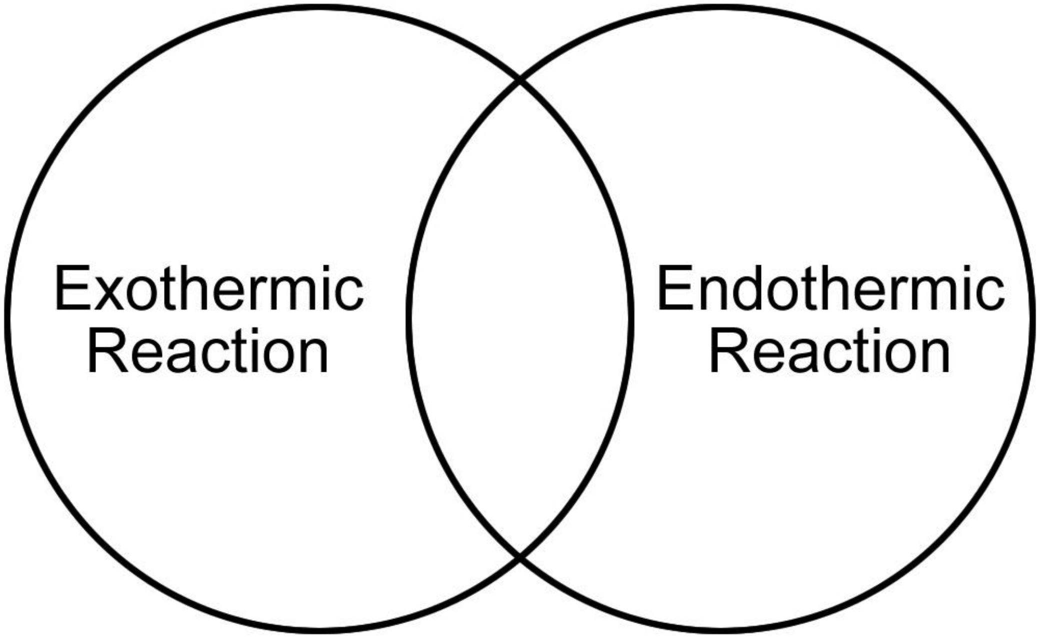
- Identify any possible sources of error in the calculations.

Higher Order Thinking Skills

In what ways are exothermic and endothermic reactions **a)** *similar* **b)** *different*?

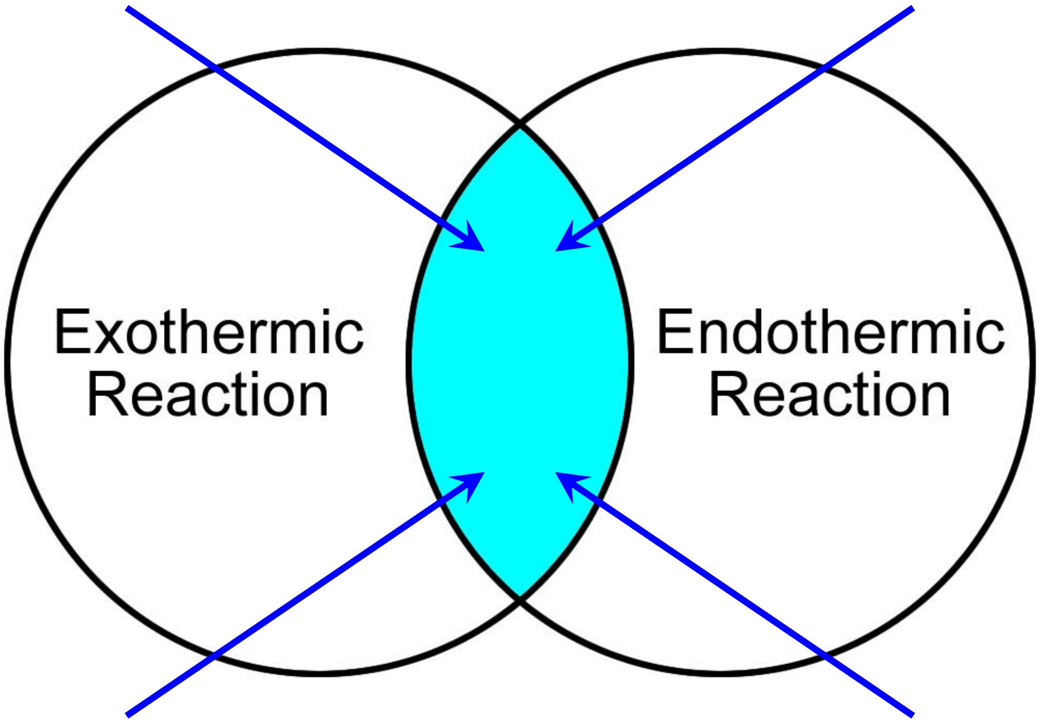
Higher Order Thinking Skills

Compare and Contrast Exothermic and Endothermic Reactions



Higher Order Thinking Skills

Compare and Contrast Exothermic and Endothermic Reactions



- Reaction involves bond formation.
- Reaction involves bond breaking.

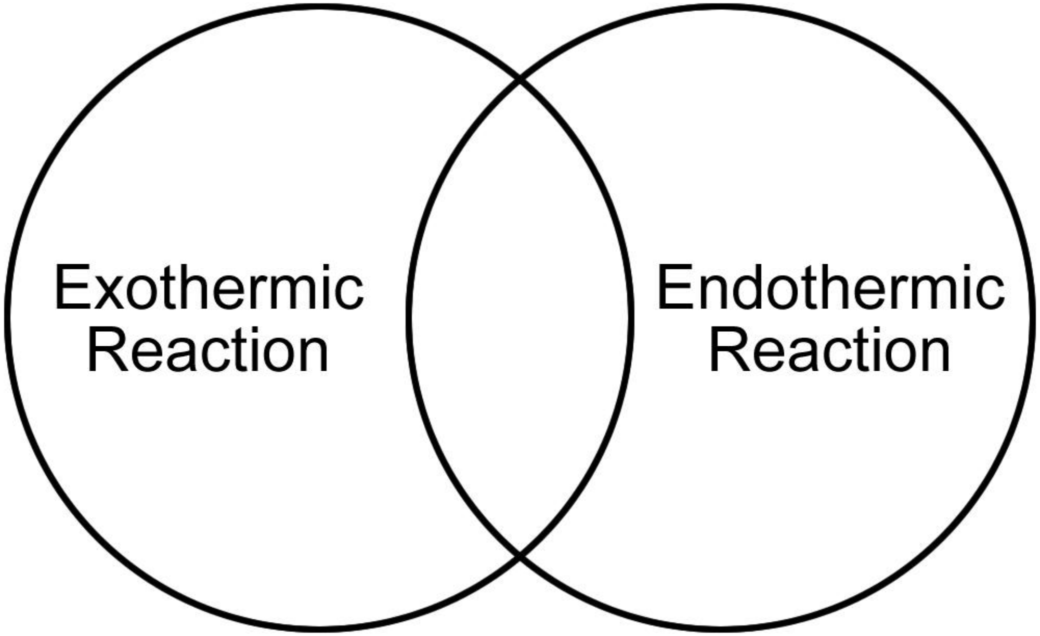
Exothermic Reaction

Endothermic Reaction

- Reaction has an activation energy.
- A catalyst will lower the activation energy.

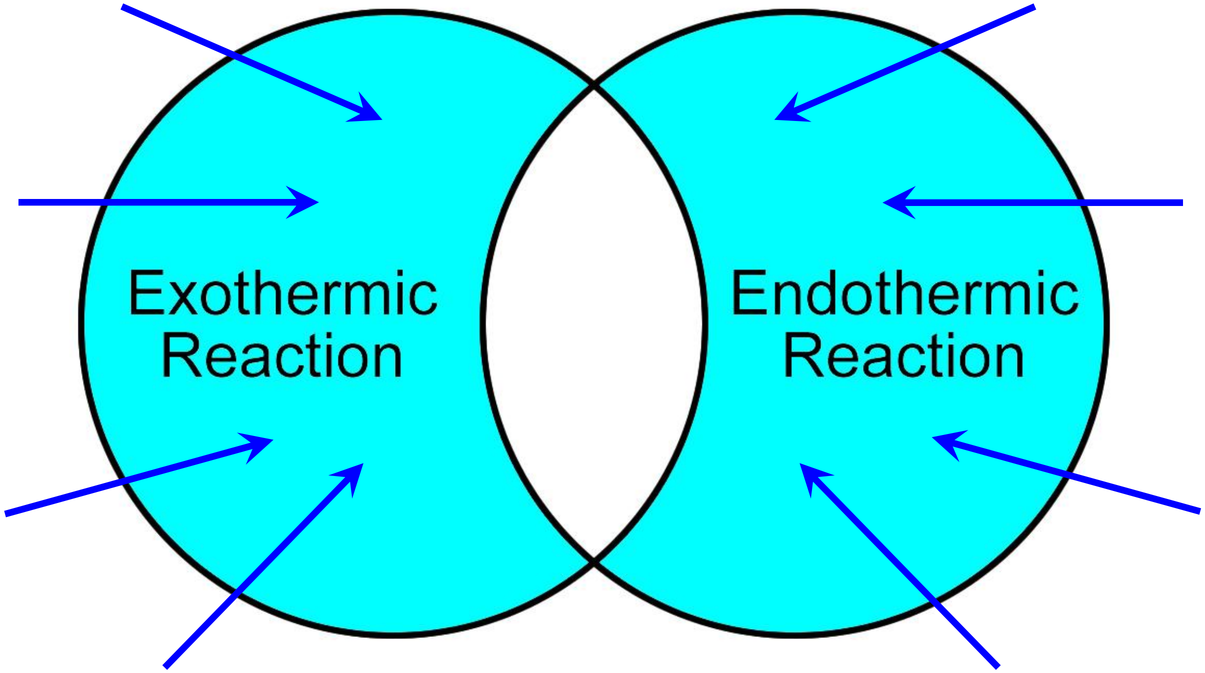
Higher Order Thinking Skills

Compare and Contrast Exothermic and Endothermic Reactions



Higher Order Thinking Skills

Compare and Contrast Exothermic and Endothermic Reactions



Exothermic Reaction

Endothermic Reaction

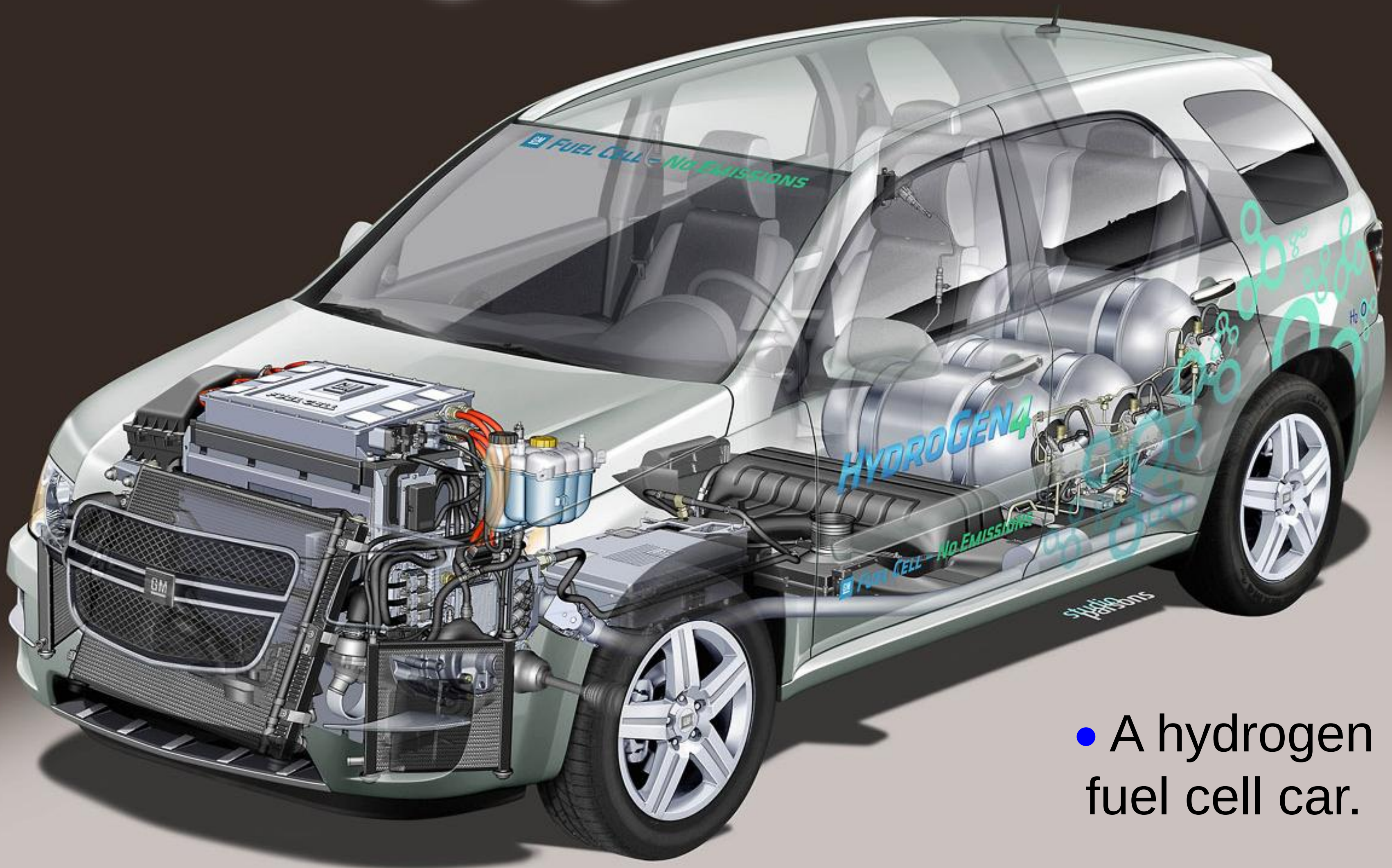
- Energy absorbed to break bonds is less than the energy released when bonds form.
- Energy of reactants is greater than energy of products.
- ΔH is negative.
- Energy released into the surroundings (temperature increases).

- 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).

Hydrogen Fuel Cells

What is a *hydrogen fuel cell*? Do they really generate electricity without causing pollution?

Hydrogen Fuel Cells

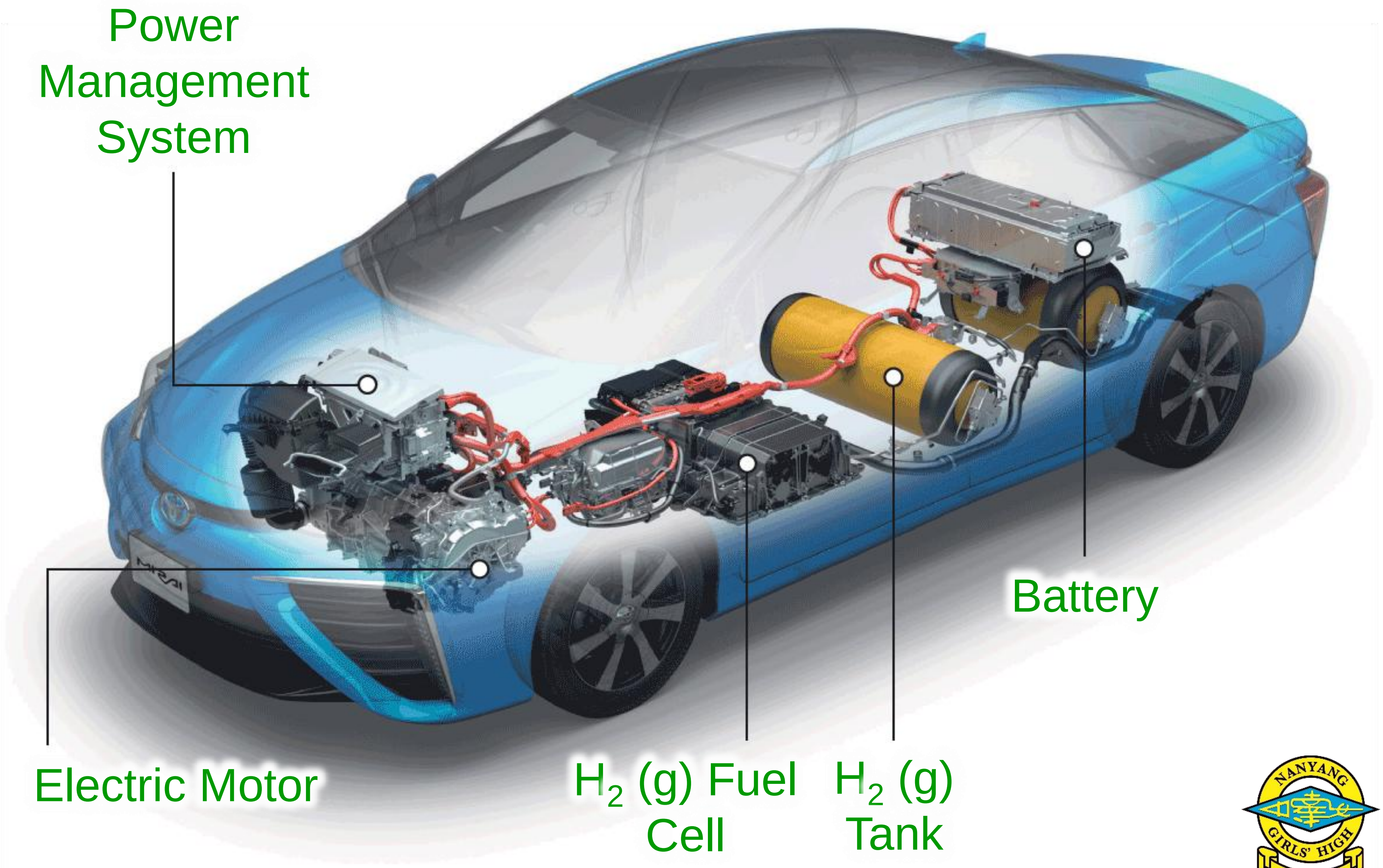


- A hydrogen fuel cell car.

Hydrogen Fuel Cells

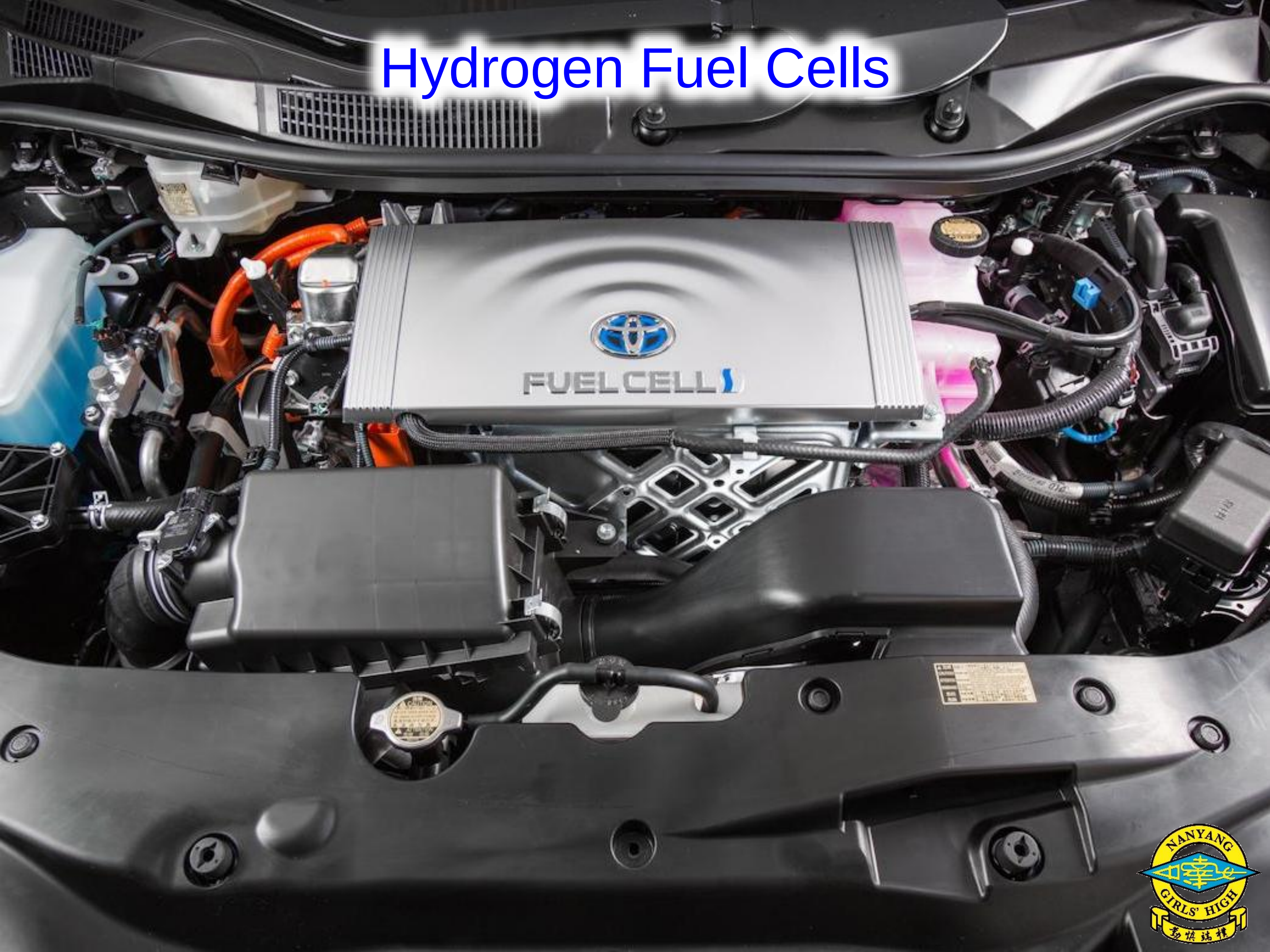
- The Toyota Mirai – the world's first commercial hydrogen fuel cell car – went on sale in Japan on 15th December 2014.

Hydrogen Fuel Cells



Hydrogen Fuel Cells

Hydrogen Fuel Cells



Hydrogen Fuel Cells

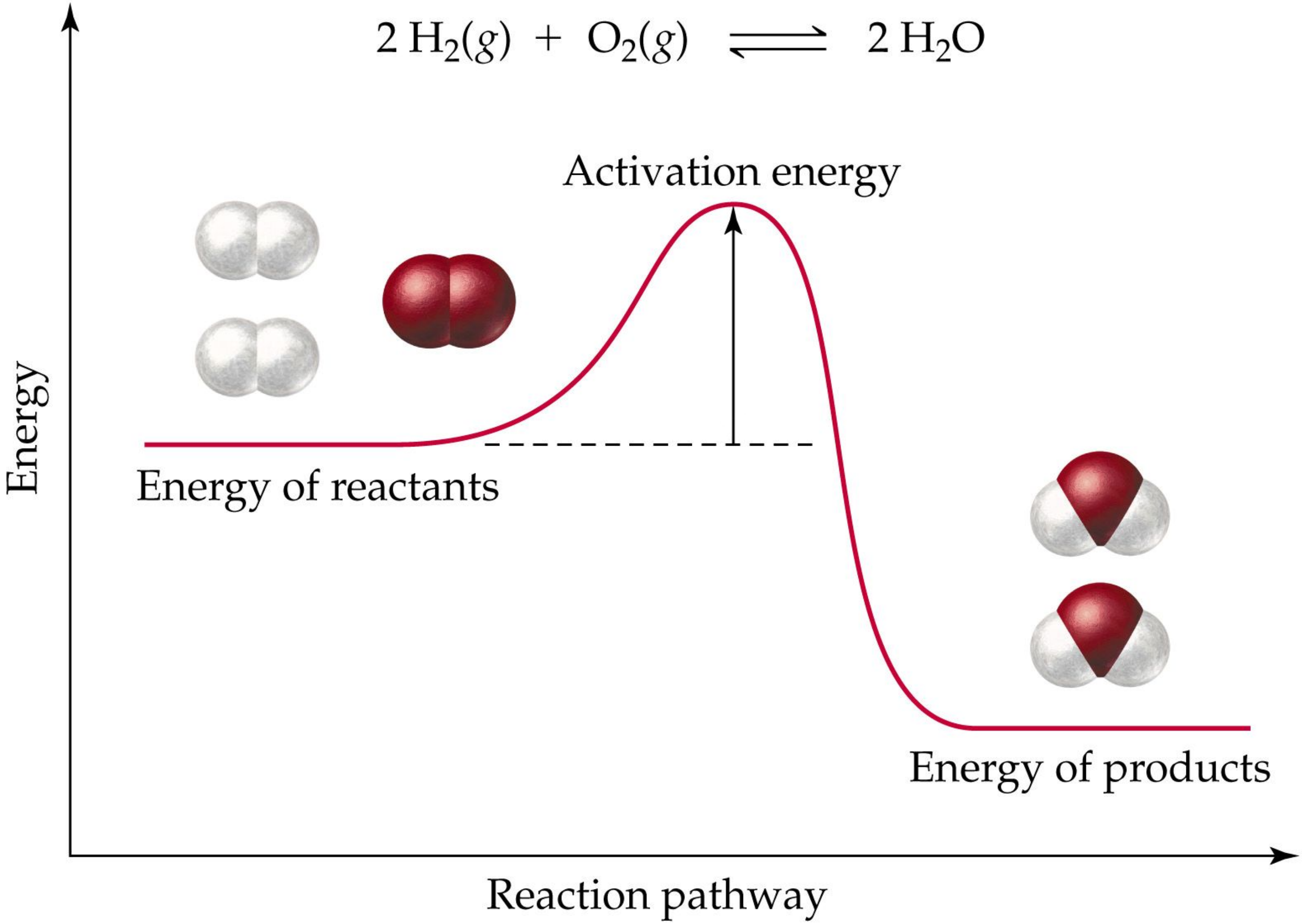
Hydrogen Fuel Cells

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

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l) \qquad \Delta H = -572 \text{ kJ mol}^{-1}$$

- 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.

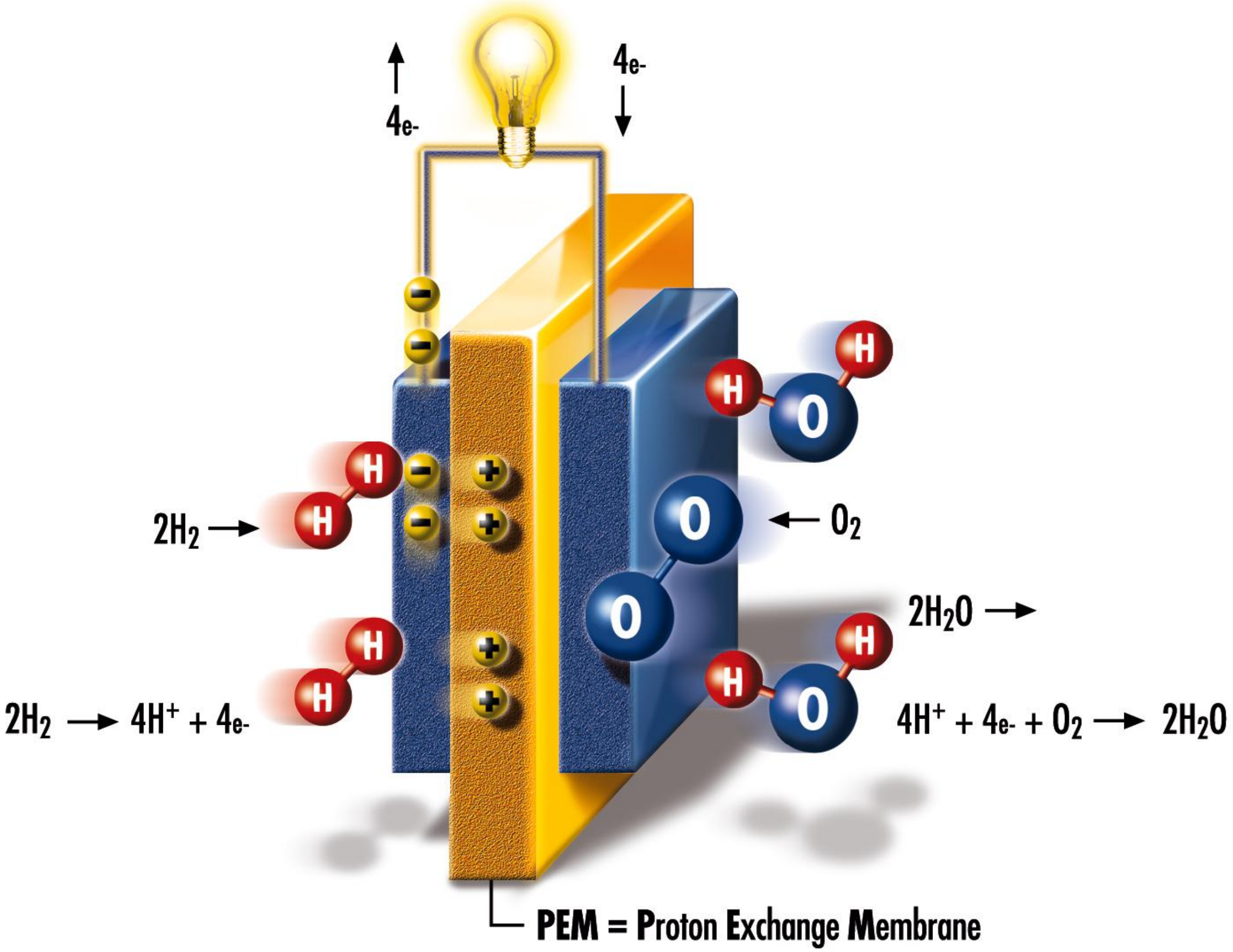
$$H_2(g) \rightarrow 2H^+(aq) + 2e^-$$

- 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.*

$$C_{20}H_{42}(s) \rightarrow C_8H_{16}(l) + C_{12}H_{24}(l) + H_2(g)$$

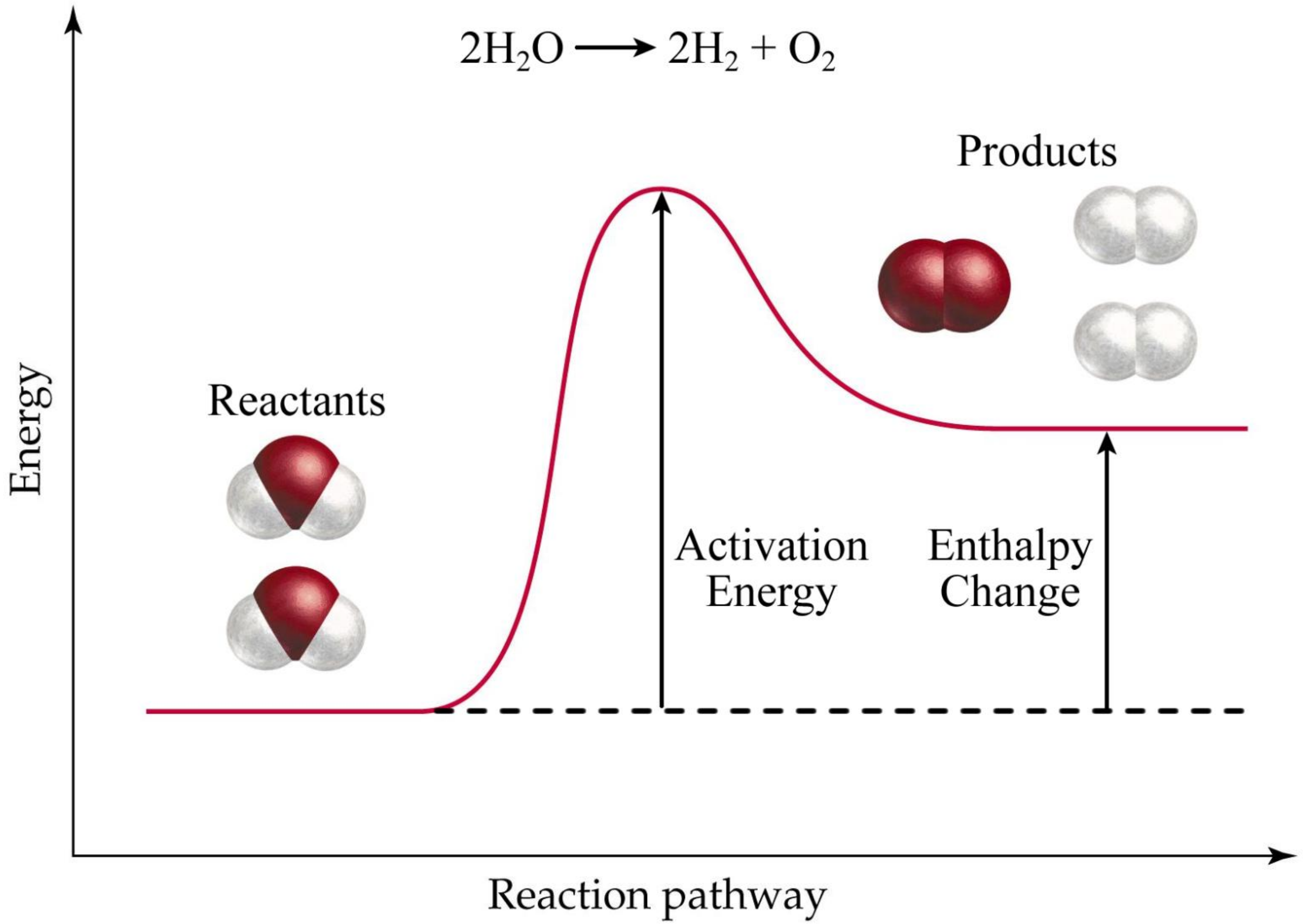
- Hydrogen can also be obtained from the *electrolysis* of water (the decomposition of water by electricity):

$$2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$$

- 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.

Hydrogen Fuel Cells

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

- The only chemical product of the reaction is H_2O, which is not considered to be a pollutant. Compare this to the chemicals that are produce by the internal combustion engine, *e.g.* CO, CO_2, NO_2 and SO_2.

Hydrogen Fuel Cells

Are hydrogen fuel cells really the solution to our *environmental problems*?

- Hydrogen is obtained by cracking of long-chain hydrocarbons and electrolysis of water, both of which require energy. If this energy is provided by *burning fossil fuels*, then the hydrogen fuel cell is not as clean as it may seem.

Experimental Design

How do I *plan an experiment* to measure the enthalpy change of a reaction?

Experimental Design

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

methanol $2CH_3OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 4H_2O(g)$

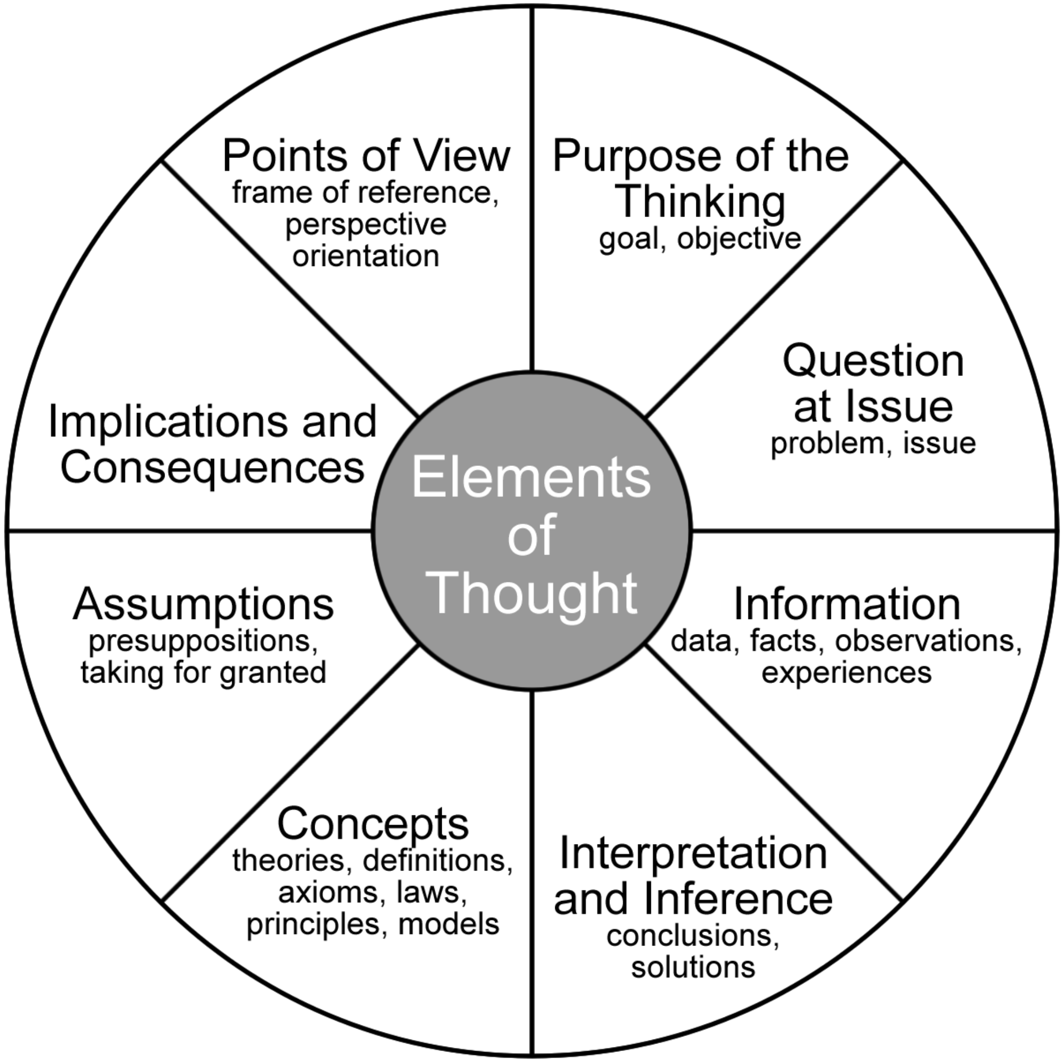
ethanol $C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$

propan-1-ol $2C_3H_7OH(l) + 9O_2(g) \rightarrow 6CO_2(g) + 8H_2O(g)$

- 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



- Experiments should be designed using critical thinking skills.

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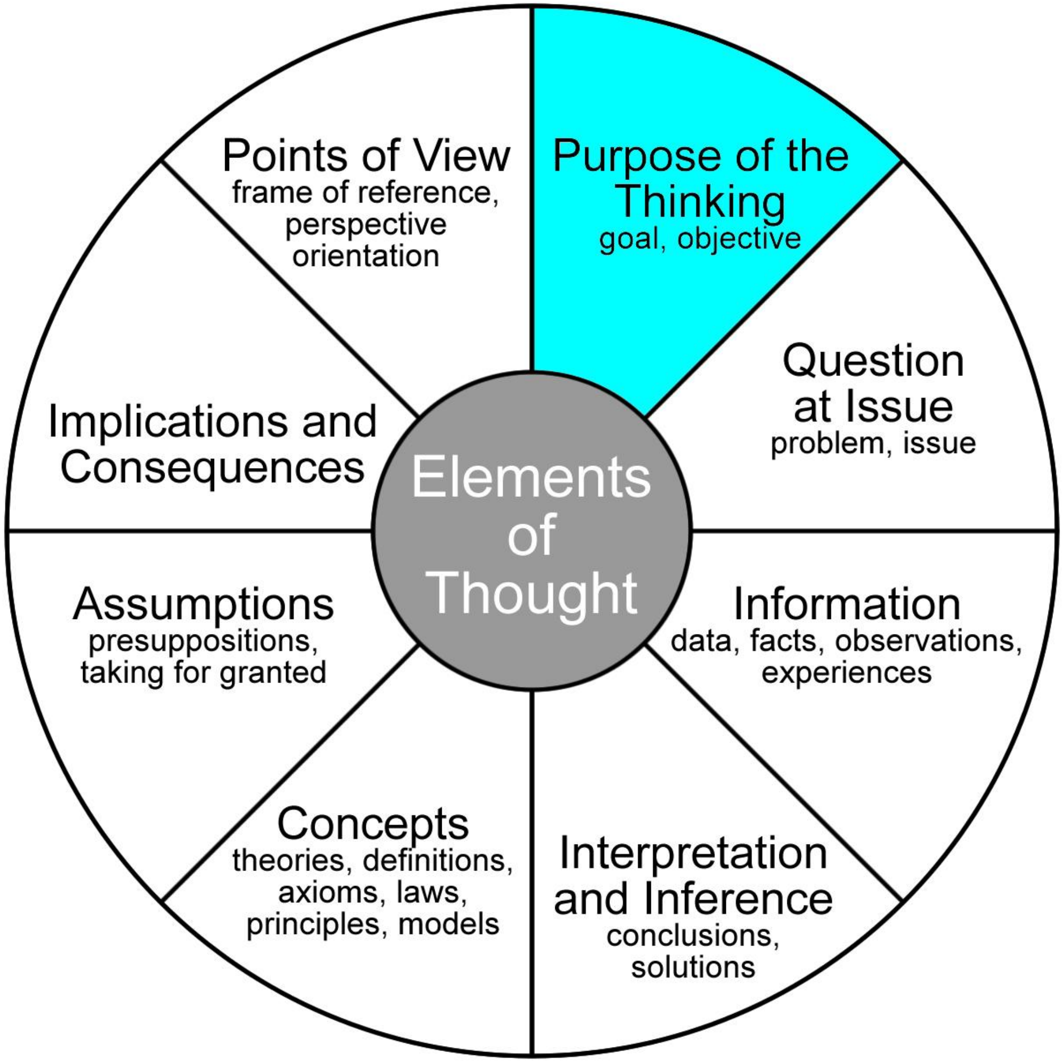

Elements of Thought

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

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

Experimental Design

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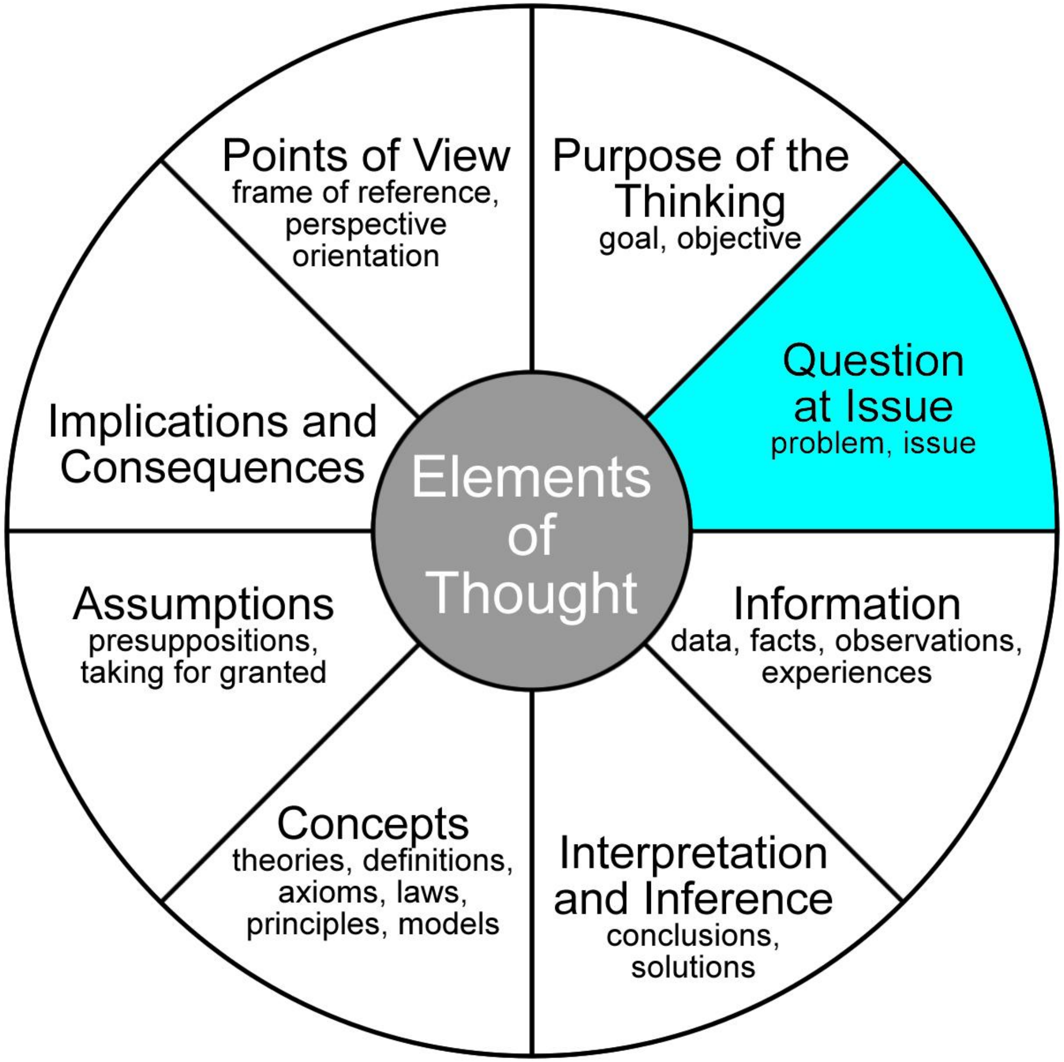

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Implications and Consequences

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

Experimental Design

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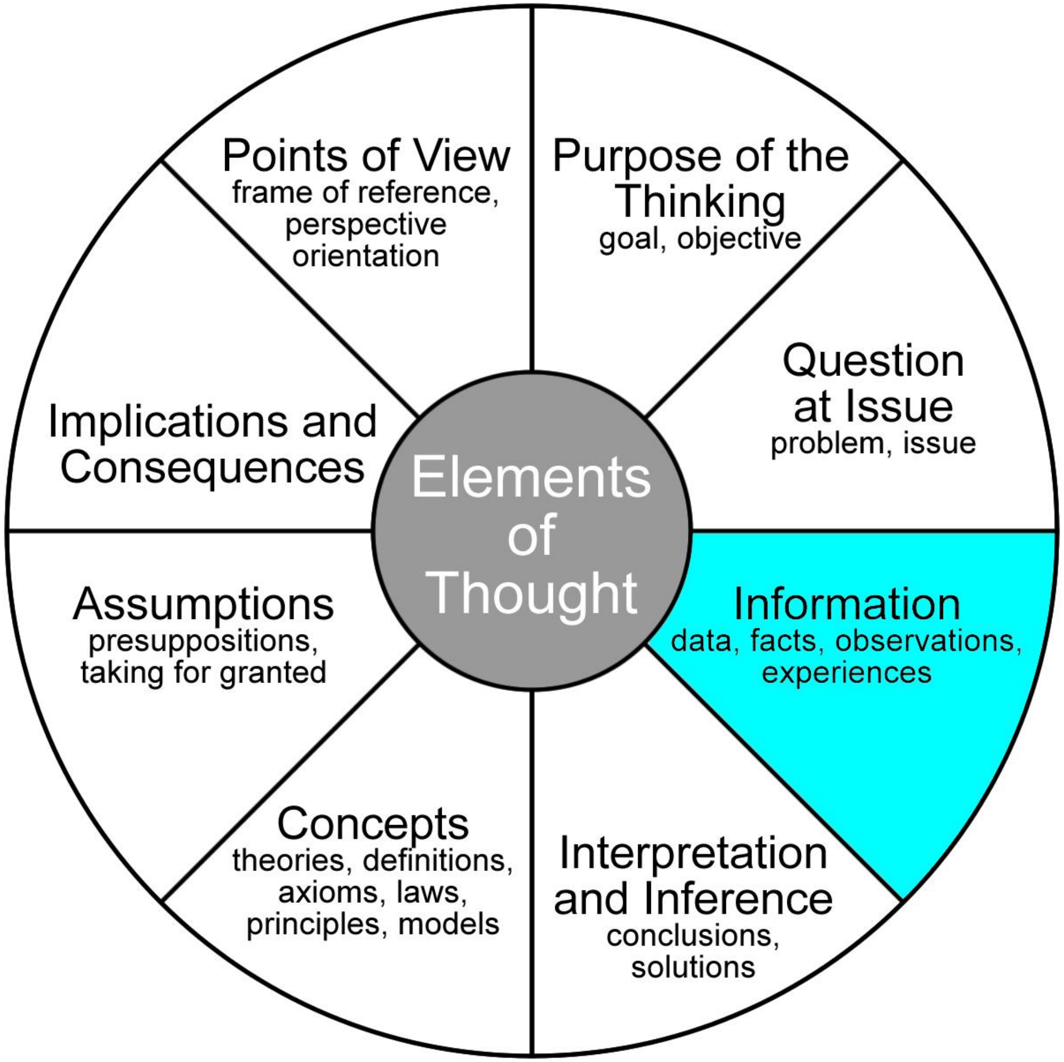

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- What essential information needs to be collected by the experiment?
- What observations need to be made?
- What are the different variables (independent, dependent and control)?

Experimental Design

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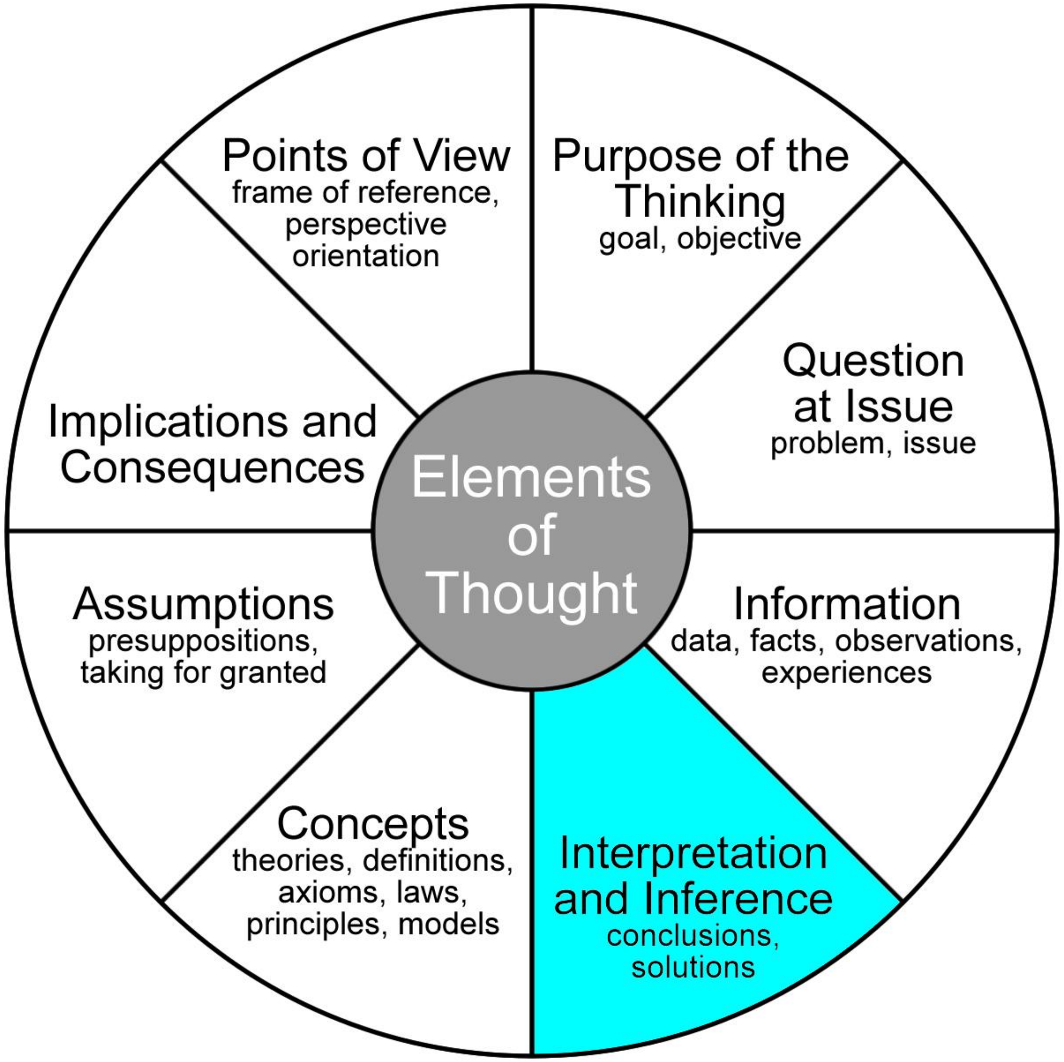

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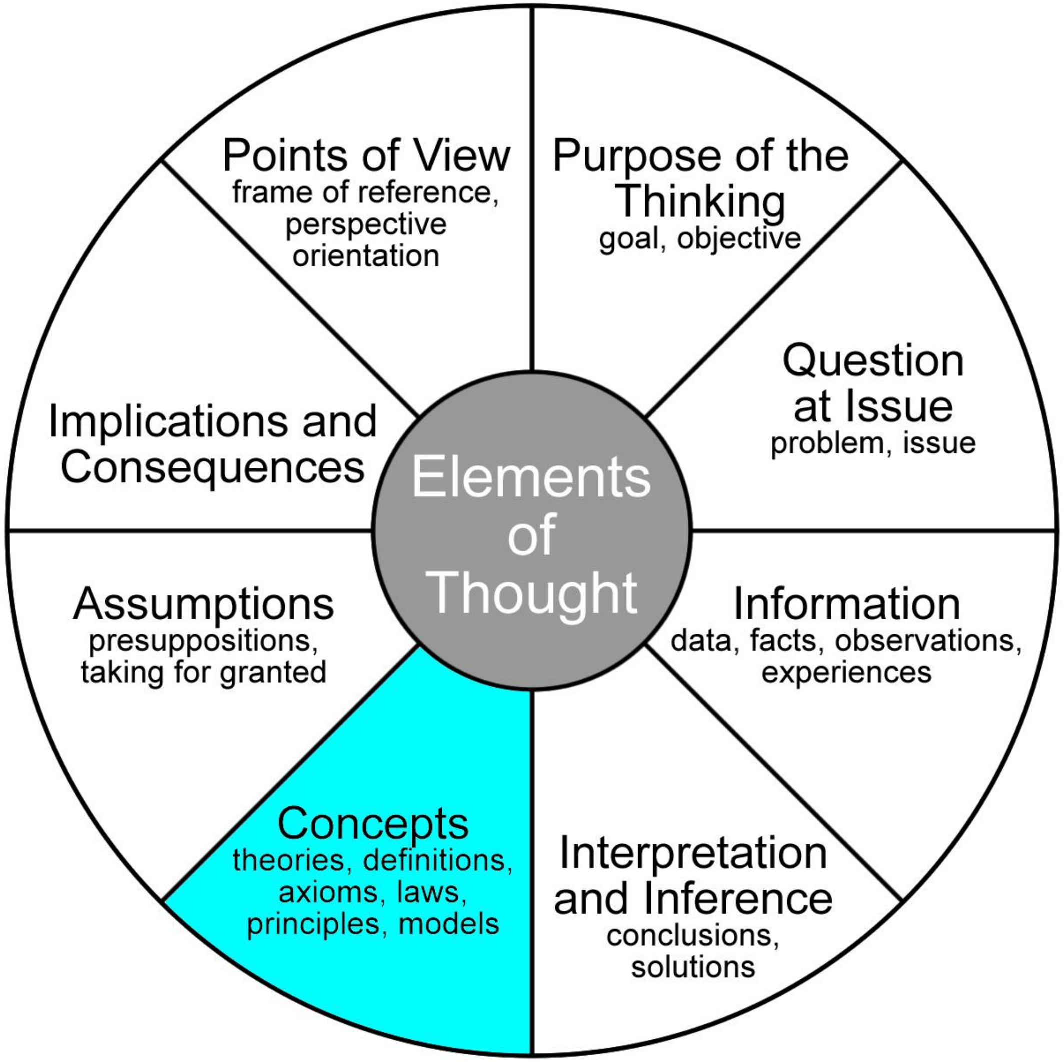
- How must the data from the experiment be manipulated and presented?
- What calculations (if any) need to be done?
- What graphs (if any) need to be plotted?

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

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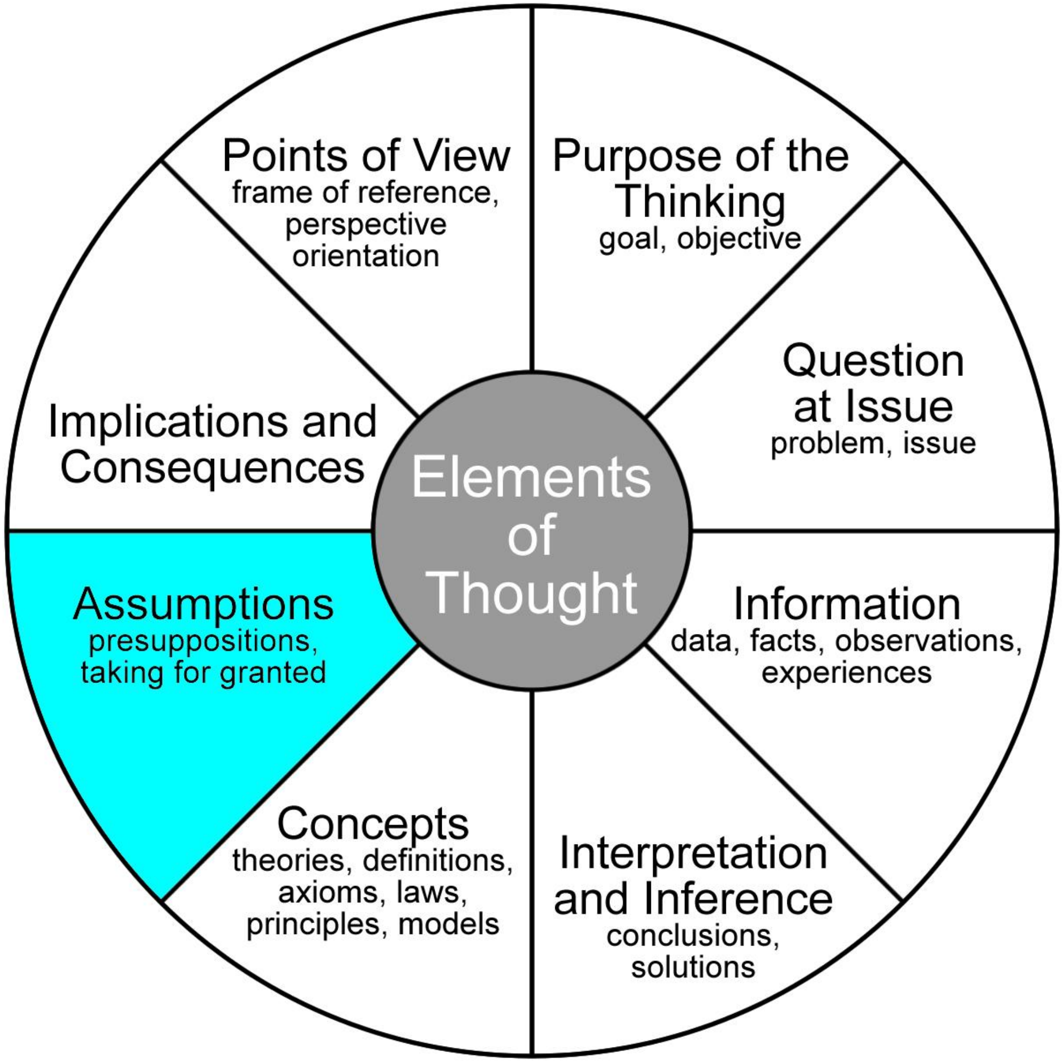

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- What assumptions are being made when designing the experiment?
- Is it reasonable to make these assumptions?
- Identifying assumptions can help to identify errors.

Experimental Design

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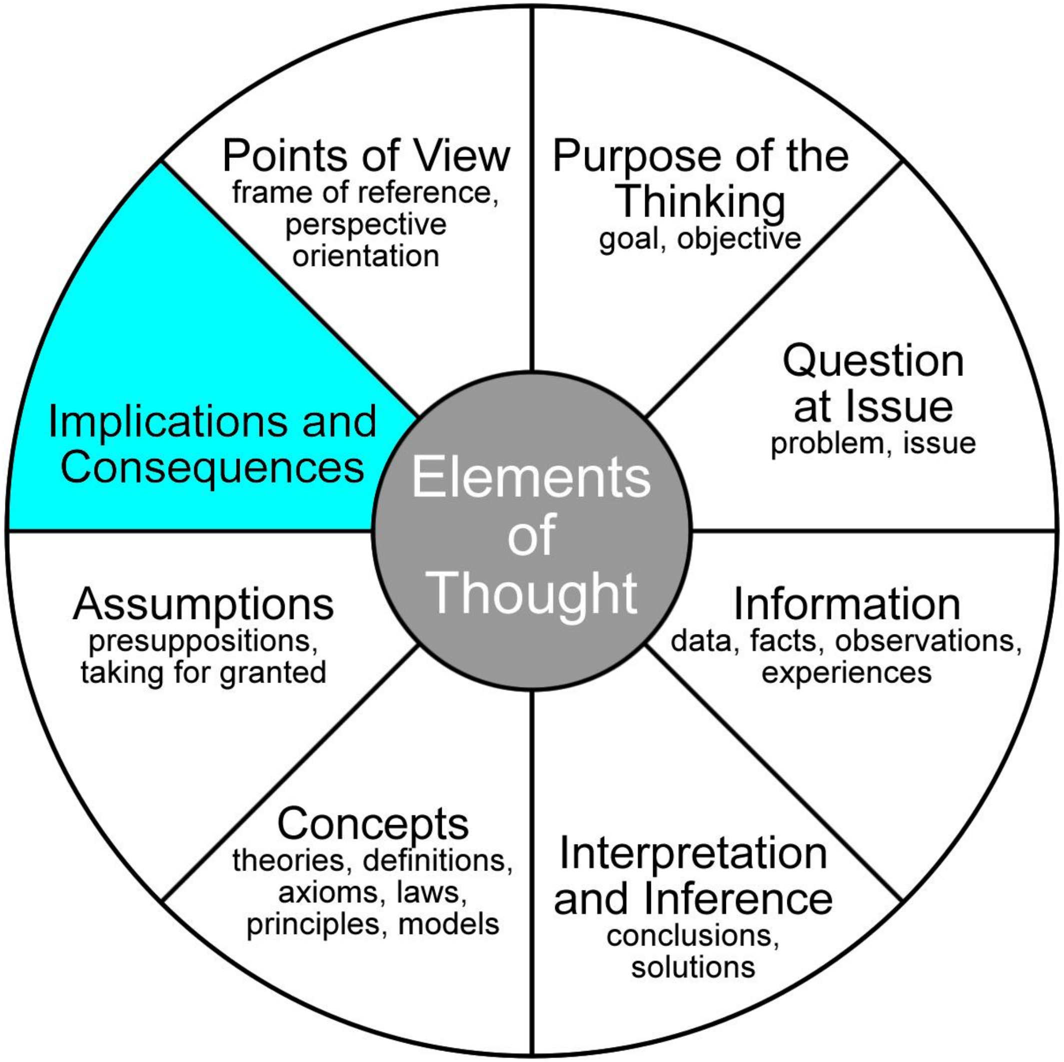

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Elements of Thought

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

Experimental Design

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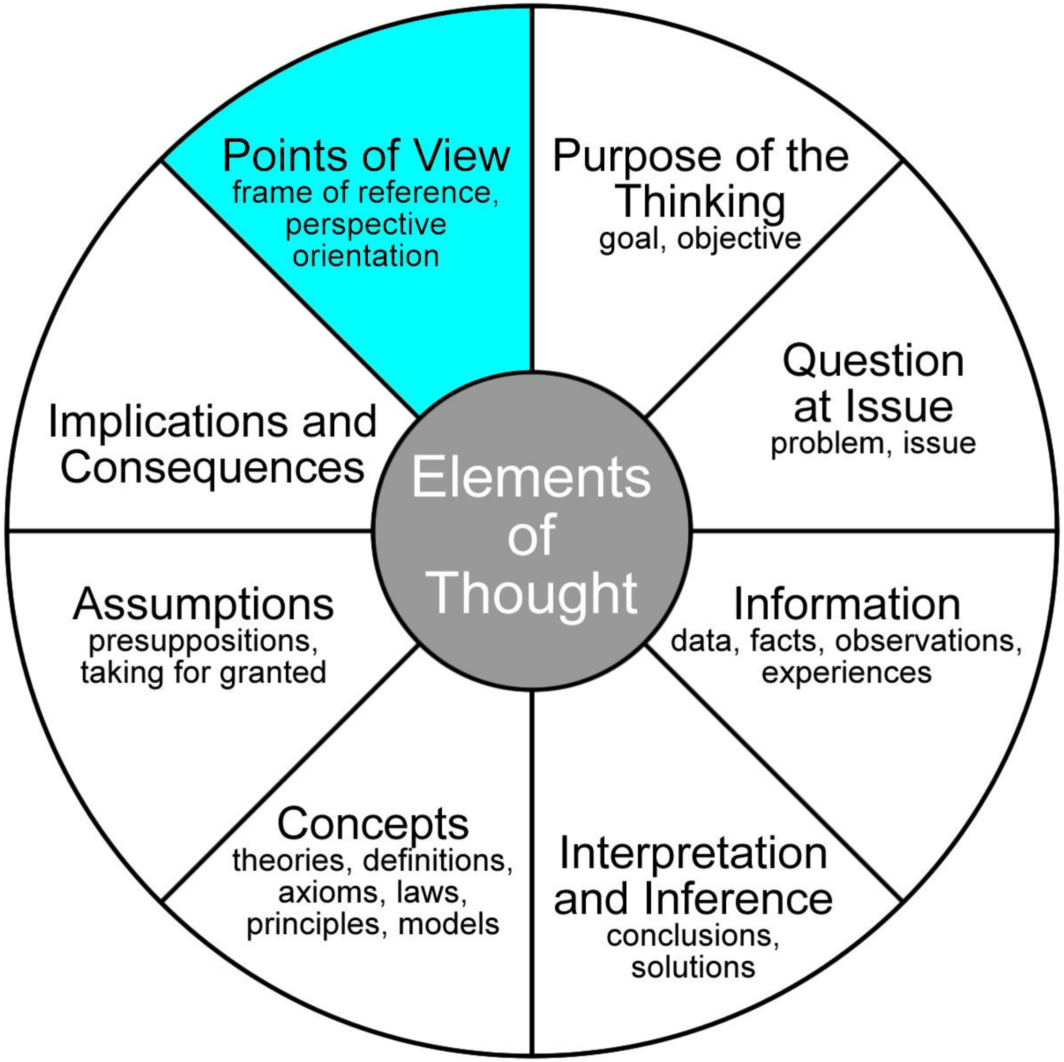

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Implications and Consequences

- 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.

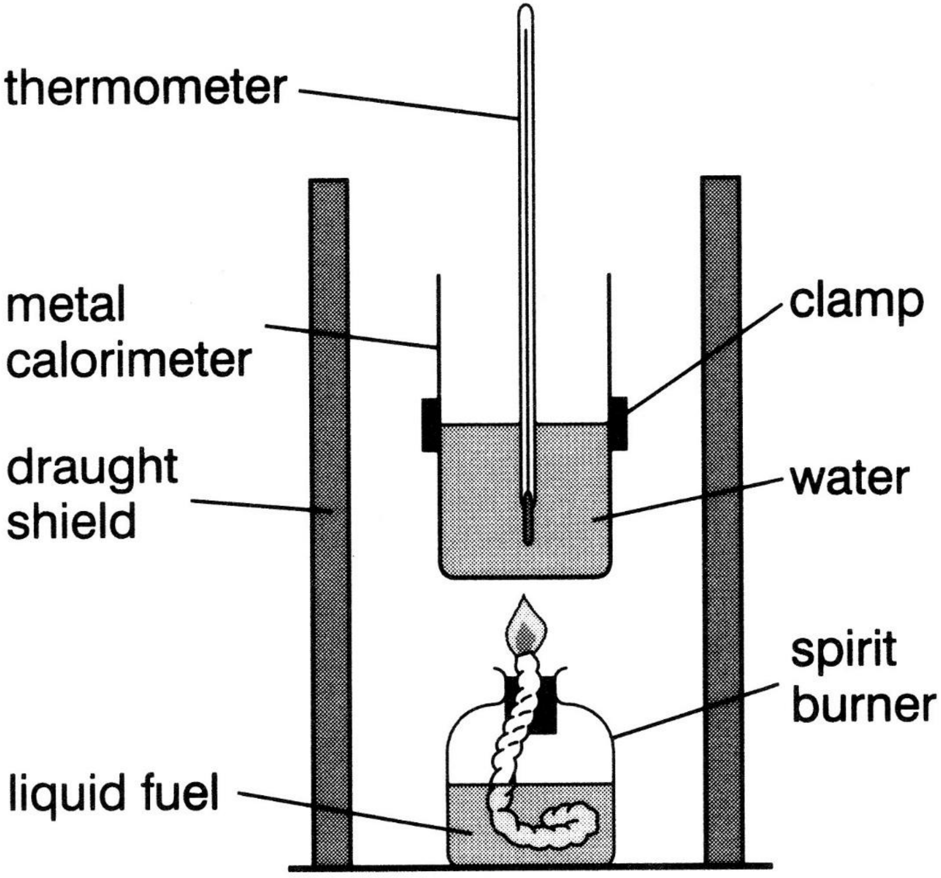
Experimental Design

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

Measured and Derived Quantities	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.

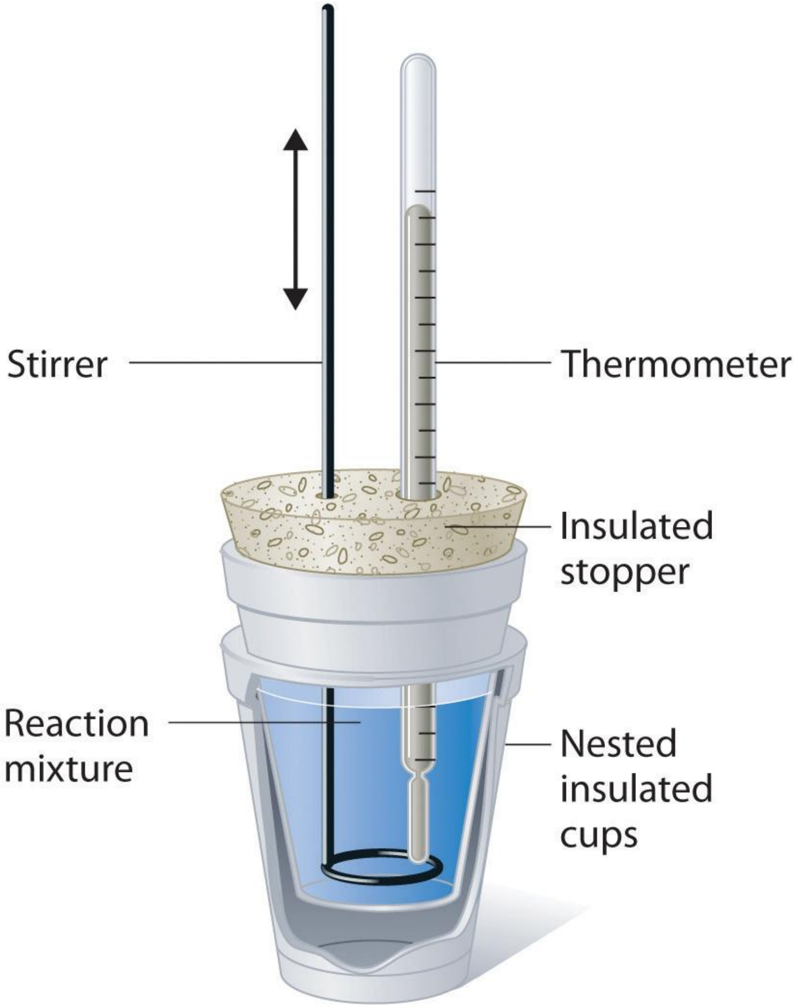
Experimental Design

- 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.

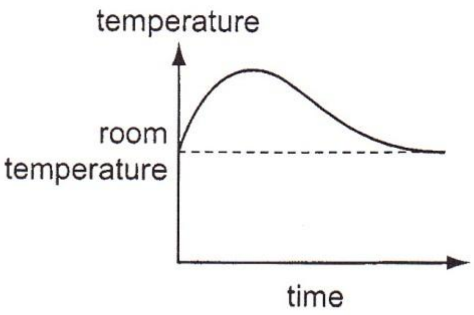
Assessment for Learning

Could I pleas have some questions to *check my understanding*?

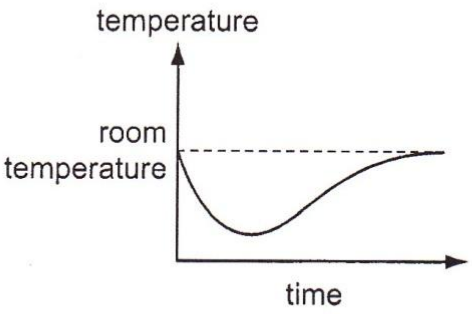
Assessment for Learning
Question 1:

- 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?

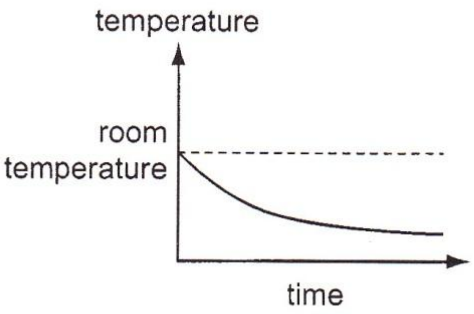
A



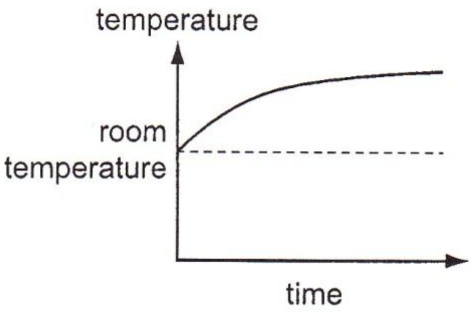
B



C



D



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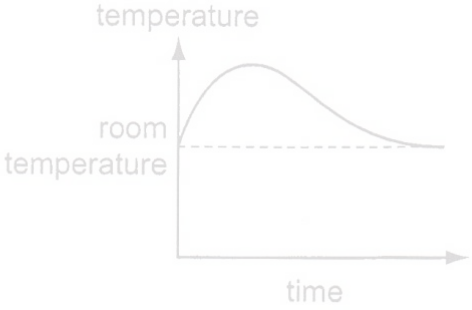

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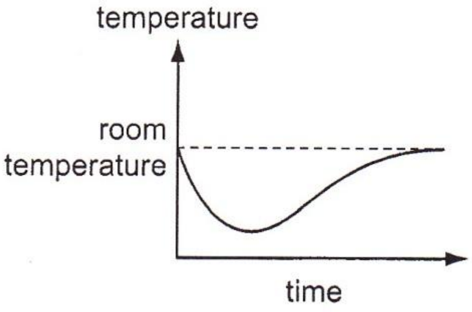

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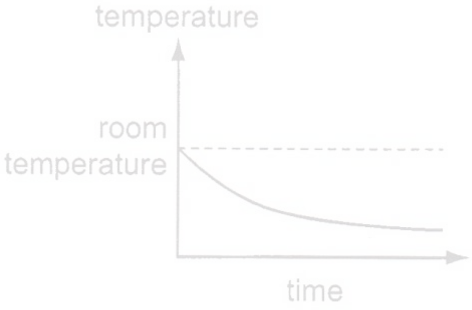

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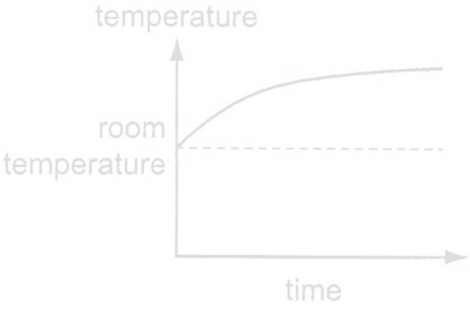

Assessment for Learning
Question 2:

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

$$\text{H–H} + \text{I–I} \rightarrow \text{H–I} + \text{H–I}$$

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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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^{-1}
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

Assessment for Learning Question 3:

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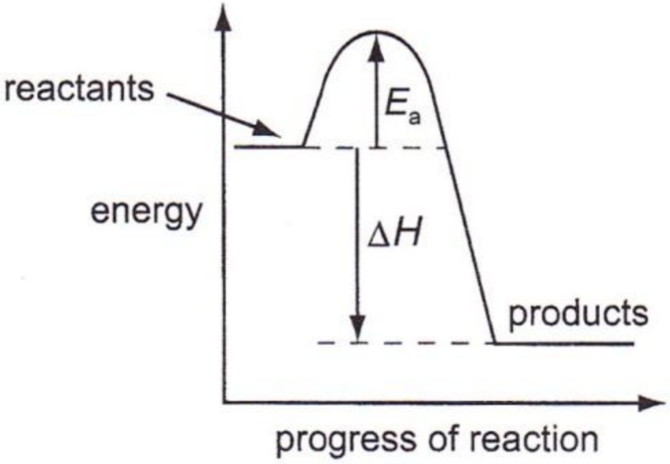
D Propane

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.

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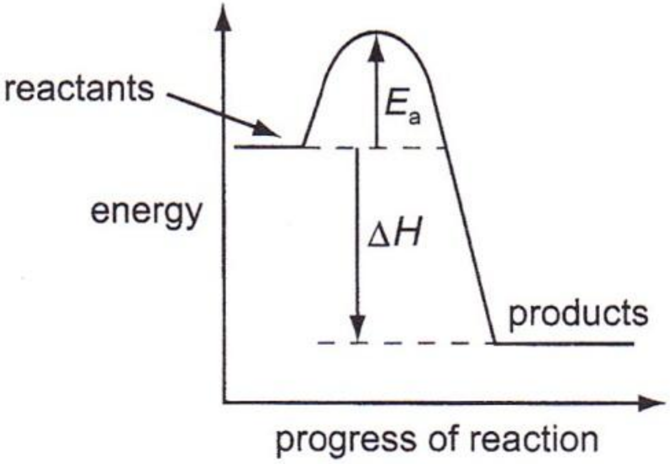

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Assessment for Learning
Question 5:

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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Energy from Chemicals – Enthalpy Changes

Presentation on

Energy from Chemicals

by Dr. Chris Slatter

stopher_john_slatter@nygh.edu.sg

Nanyang Girls’ High School
2 Linden Drive
Singapore
288683

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