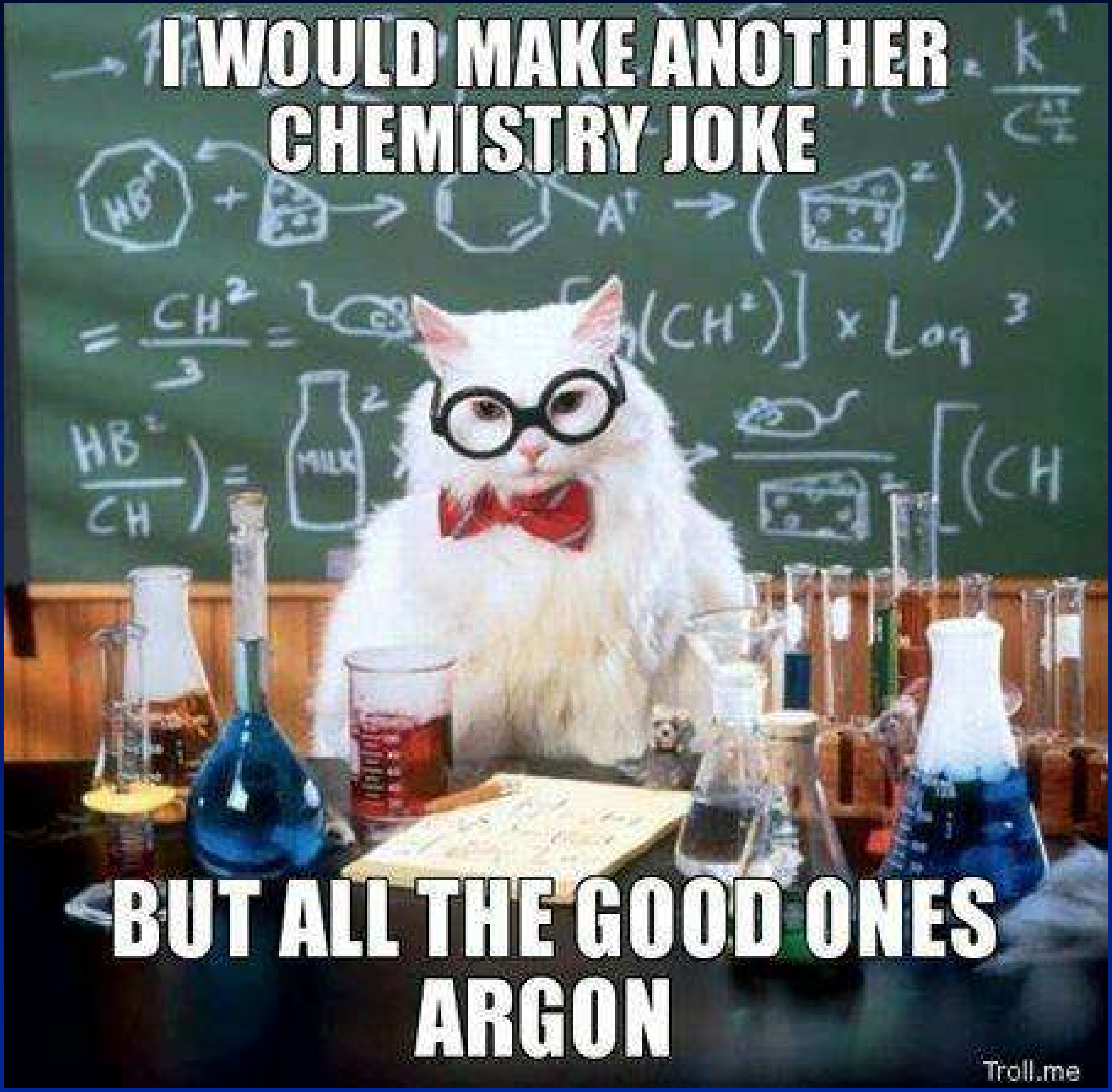
Chapter 10

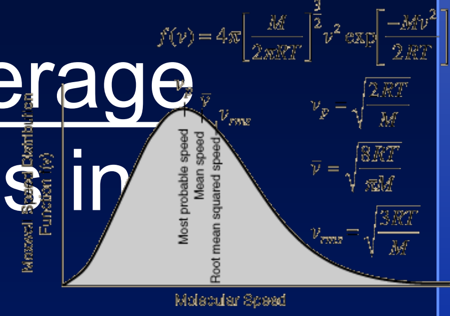
Causes of Change

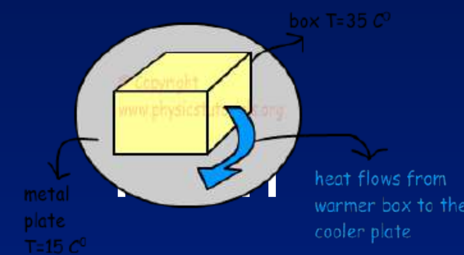

I WOULD MAKE ANOTHER CHEMISTRY JOKE
BUT ALL THE GOOD ONES ARGON
Troll.me

10.1 Energy Transfer

- Define enthalpy
- Distinguish between heat and temperature
- Perform calculations using molar heat capacity.
- Perform calculations using specific heat capacity.

Temperature is different than heat

- 

Temperature measures the <u>average kinetic energy</u> of the particles in
sample of matter.
- Memorize!!!
- 

Heat - the energy transferred
hot things to cold things.
- A drop of boiling water hurts,
- kilogram of boiling water kills (same temperature, but more heat)

Enthalpy

- **Enthalpy** - total energy content of something.
- Old measurement is calories (food)
- Now we use **joules**
- Example - When 1 gram of water absorbs 4.184 joules of energy (heat) its temperature goes up 1 º

Units of heat

- calories or Joules
- 1 calorie = amount of heat needed to raise the temperature of 1 gram of water by 1ºC
- It takes 4.184 J to do the same thing.
- So, 1 calorie = 4.184 J
- a food Calorie is really a <u>kilo</u>calorie

Energy conversions

- How much energy is needed to heat 15 g of water by 25ºC? We'll do in J & cal. . .
- Q=m(Δt)Cp
  - 15 g • 25ºC • 4.184 J/gºC = 1569 J
  - 15 g • 25ºC • 1 cal/gºC = 375 cal
- Convert 10 cal to J . . .
  - 10 cal • 4.184 J/cal = 41.84 J
- Convert 10 J to cal . . .
  - 10 J • 1 cal/4.184 J = 2.39 cal
- Use these to help with online HW.

Some Equalities for Heat

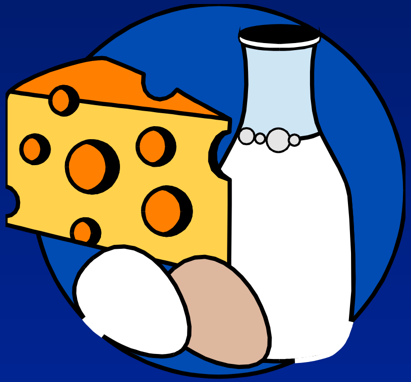
Heat is measured in calories or joules

- 1 kcal = 1000 cal = 1 Cal (a food calorie)
- 1 calorie = 4.184 J
- 1 kJ = 1000 J

Energy and Nutrition

1 Calorie (nutritional) = 1 kcal

1 Cal= 1000 cal



Caloric Food Values

Carbohydrate = 4 kcal/g

Fat = 9 kcal/g

ein = 4 kcal/g

Foods and Calories

FoodCarbo	Fat	Protein	Energy(kcal
carrots, 1 cup 11	0	1	50
banana 26	0	1	110
egg 0	6	6	80
chicken (no skin) 0	3	20	110
beef (3 oz) 0	5	22	130

Learning Check

1.0 cup of whole milk contains 12 g of carbohydrate, 9.0 g of fat, and 9.0 g of protein. How many kcal (Cal) are obtained? Answer? . . .

1) 48 kcal

2) 81 kcal

3) 165 kcal

*** Steps follow

Solution

3) 165 kcal = 165 Calories (capital "C")

12 g carbo x 4 kcal/g = 48 kcal

9.0 g fat x 9 kcal/g =81 kcal

9.0 g protein x 4 kcal/g=36 kcal

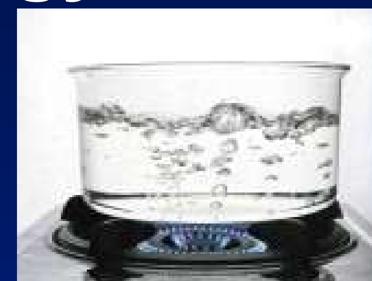
Total kcal= 165 kcal

Heat Capacity & Specific Heat

- Why do some foods stay hot longer than others?
- Why is the beach sand hot, but the water is cool on the same hot day?

Specific Heat

Different substances have different capacities for storing energy

It may take 20 minutes to  er to 75°C.

However, the same mass of aluminum  might require 5 minutes a e same amount of copper m ke only 2 minutes to reach the same temperature.

Learning Check

When you heat 200 g of water for 1 minute, the water temperature rises from 10°C to 18°C.

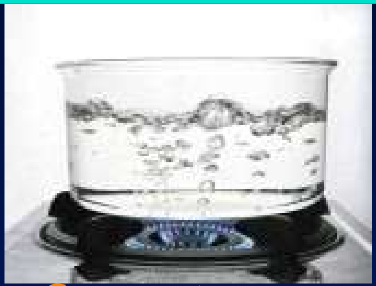
200 g

400 g

If you heat 400 g of water at 10°C in the same pan with the same amount of heat for 1 minute, what would you expect the final temperature to be?

1) 10 °C 2) 14°C 3) 18°C

Solution



2) 14°C

Heating twice the mass of water using the same amount of heat will raise the temperature only half as much.

200 g

400 g

Specific Heat Values

Specific heat is the amount of heat needed to raise the temperature of 1 g of a substance by 1°C

	cal/g°C	J/g°C
water	1.00	4.184
aluminum	0.22	0.90
copper	0.093	0.39
silver	0.057	0.24
gold	0.031	0.13

Learning Check

A. A substance with a large specific heat 1) heats up quickly 2) heats up slowly

2) heats up slowly

B. When ocean water cools, the surrounding air 1) cools 2) warms 3) stays the same

2) warms

C. Sand in the desert is hot in the day, and cool at night. Sand must have a 1) high specific heat 2) low specific heat

2) low specific heat

Some things heat up easily

- Some take a great deal of energy to change their temperature.
- Specific Heat Capacity - amount of heat to change the temperature of 1 g of a substance by 1ºC

Molar & Specific Heat

- Table of molar heat on p. 343. Also table of specific heat in these slides.
- Water has a high specific heat
- 75.3 J/K•mol or 4.184 J/gºC
- Check your units!
- The amount of heat it takes to heat something is the same as the amount of heat it gives off when it cools because...
- Law of conservation of energy.

Table 1 Molar Heat Capacities of Elements and Compounds

Element	C (J/K•mol)	Compound	C (J/K•mol)
Aluminum, Al(*s*)	24.2	Aluminum chloride, $AlCl_3(s)$	92.0
Argon, Ar(*g*)	20.8	Barium chloride, $BaCl_2(s)$	75.1
Helium, He(*g*)	20.8	Cesium iodide, CsI(*s*)	51.8
Iron, Fe(*s*)	25.1	Octane, $C_8H_{18}(l)$	254.0
Mercury, Hg(*l*)	27.8	Sodium chloride, NaCl(*s*)	50.5
Nitrogen, $N_2(g)$	29.1	Water, $H_2O(g)$	36.8
Silver, Ag(*s*)	25.3	Water, $H_2O(l)$	75.3
Tungsten W(*s*)	24.2	Water, $H_2O(s)$	37.4

TABLE 6.1 Specific Heat Values for Some Elements, Compounds, and Common Solids

Substance	Name	Specific Heat (J/g · K)
Elements		
Al	Aluminum	0.902
C	Graphite	0.720
Fe	Iron	0.451
Cu	Copper	0.385
Au	Gold	0.128
Compounds		
$NH_3(\ell)$	Ammonia	4.70
$H_2O(\ell)$	Water—liquid	4.184
$C_2H_5OH(\ell)$	Ethanol	2.46
$(CH_2OH)_2(\ell)$	Ethylene glycol (antifreeze)	2.42
$H_2O(s)$	Water—ice	2.06
$CCl_4(\ell)$	Carbon tetrachloride	0.861
$CCl_2F_2(g)$	Dichlorodifluoromethane (a chlorofluorocarbon)	0.598
Common Solids		
Wood		1.76
Cement		0.88
Glass		0.84
Granite		0.79

Molar Heat Capacity

- The amount of heat necessary to raise the temperature of 1 mole of the substance 1 K
- Every substance has its own special value.
- Abbreviated as Cp
- Molar Heat Capacity of Water = 75.3 J/mol•K

Specific Heat Capacity

- The amount of heat necessary to raise the temperature of 1 gram of the substance 1°C
- Every substance has its own special value
- Cp
- Specific Heat Capacity of Water = 4.184 J/g°C

Molar Heat

$q = n \cdot \Delta T \cdot Cp$

whereq => heat, J

n => moles

ΔT = change in temperature, K

Cp => molar heat, J/mol•K

Specific Heat

$q = m \bullet \Delta t \bullet Cp$

whereq => heat, J

m => mass, g

Δt = change in temperature, ^oC

Cp => specific heat, $J/g \bullet {}^oC$

Be sure to check if you are doing the problem in grams or moles! The Cp values will be different.

Molar Heat Calculations

A hot-water bottle contains 41.6 moles of water at 338 K. If the water cools to body temperature (310 K), how many joules of heat could be transferred to sore muscles? q = n•ΔT•Cp(1st calculate ΔT)

Heat = q = n x ΔT x Cp (H_2O)

41.6 mol x 28 K x $\frac{75.3\text{ J}}{\text{mol•K}}$

= 88 000 J (remember significant figures)

= 88 kJ (since 1 kJ/1000 J)

Specific Heat Calculations

A hot-water bottle contains 750 g of water at 65°C. If the water cools to body temperature (37°C), how many joules of heat could be transferred to sore muscles? q = m•Δt•Cp

Heat = q = m x ΔT x Cp (H_2O)

750 g x 28°C x 4.184 J

= 88 000 J (remember significant figures)

= 88 kJ (since 1 kJ/1000 J)

Learning Check – do as class

How many kilojoules are needed to raise the temperature of 120 g of water from 15°C to 75°C? . . .

Answer is 30. kJ Calculation . . .

120 g x (75°C - 15°C) x $\frac{4.184 \text{ J}}{\text{g°C}}$ x $\frac{1 \text{ kJ}}{1000 \text{ J}}$

Now You Try It

- It takes 1950 joules to heat 10.0 mol of a metal from 295 K to 302 K. What is its molar heat capacity? What metal is it? Hints: use $q = n \cdot \Delta T \cdot C_p$, solve for C_p, then use table on p. 343.Answer? . . .
- Cp = 27.8 J/mol•K = Mercury
- $q = n \cdot \Delta T \cdot C_p$ so $C_p = q/(n \cdot \Delta T)$ C_p = (1950 J)/(10.0 mol • 7.0 K) = **27.8 J/mol•K (Mercury)**

Now You Try It

- Iron has a specific heat of 0.449 J/gºC. How much heat will it take to change the temperature of 48.3 g of iron by 32.4ºC? Answer . . .?
- q = 703 J
- q = m•ΔT•C_p = (48.3 g)(32.4ºC)(0 J/gºC)= 703 J

Heat Transfer pp

$q_{lost} = - q_{gained}$

$(m \cdot \Delta t \cdot Cp)_{lost} = - (m \cdot \Delta t \cdot Cp)_{gained}$

You will use this principle in lab A16

In the next slide we solve for $t_{final\ of\ both}$

In lab A16 you will solve for $t_{initial\ of\ metal}$

See LD 3: 17.1 Hot metal in water

Just follow the example, don't take notes

If 100. g of iron at 100.0°C is placed in 200. g of water at 20.0°C in an insulated container, what will be the final temperature, °C, of the iron & water when both are at the same temperature? Iron specific heat is 0.106 cal/g°C.

$(100.\text{g} \bullet 0.106\text{cal/g}^{o}\text{C} \bullet (T_f - 100.)^{o}\text{C}) = q_{lost}$

$- q_{gained} = (200.\text{g} \bullet 1.00\text{cal/g}^{o}\text{C} \bullet (T_f - 20.0)^{o}\text{C})$

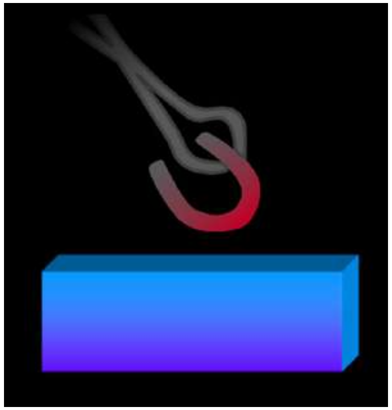
$[10.6(T_f - 100.^{o}\text{C})]_{lost} = [- 200.(T_f - 20.0^{o}\text{C})]_{gained}$

$[10.6T_f - 1060^{o}\text{C}]_{lost} = [- 200.T_f + 4000^{o}\text{C}]_{gained}$

Collect like terms . . .

$(10.6 + 200.)T_f = (1060 + 4000)^{o}\text{C}$

$T_f = (5060/211.)^{o}\text{C} = 24.0^{o}\text{C}$



10.2 Using Enthalpy

- Define thermodynamics
- Understand thermodynamic equations as being endothermic or exothermic

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Energy

- Energy is measured in Joules
- Every reaction has an energy ch

- Exothermic reactions release ene usually in the form of heat.
- Endothermic reactions absorb energy

- Energy is stored in bonds betw ns
- Sci 13 & Sci 15

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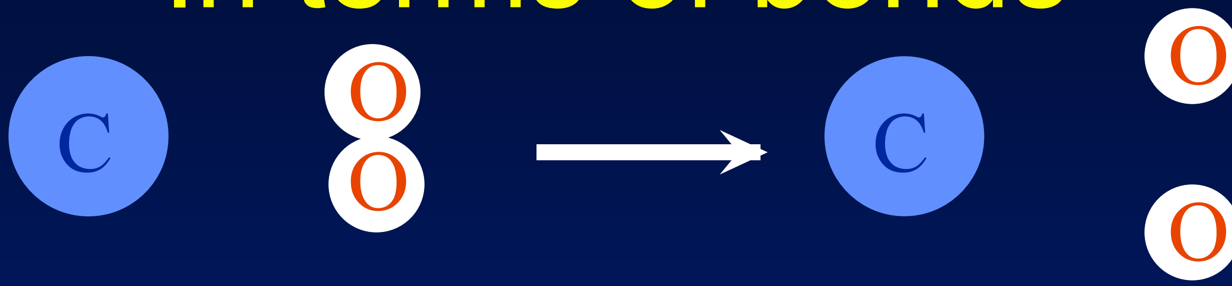
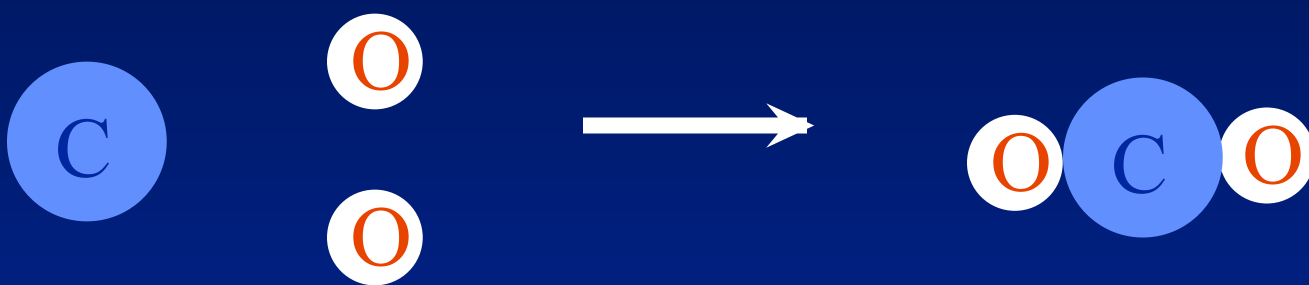
In terms of bonds



<u>Breaking</u> this bond <u>requires</u> energy



<u>Making</u> these bonds <u>releases</u> energy

In this case making the bonds gives more energy than breaking them, (so exothermic).

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Chemistry Happens in

- MOLES
- An equation that includes energy is called a thermochemical equation
- $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O + 802.2\ kJ$
- 1 mole of CH_4 makes 802.2 kJ of energy.
- When you make 802.2 kJ of energy you also make 2 moles of water

Exothermic *vs.* Endothermic

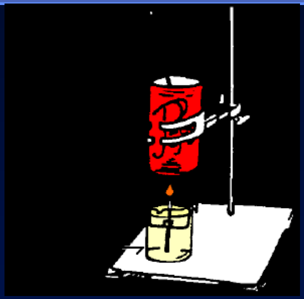
- $CH_4 + 2O_2 \rightarrow CO_2 + 2\ H_2O + 802.2\ kJ$

  This is exothermic because it makes energy. $\Delta H = -802.2\ kJ$

- The opposite reaction is endothermic.
- $CO_2 + 2H_2O + 802.2\ kJ \rightarrow CH_4 + 2O_2$

  $\Delta H = +802.2\ kJ$

Enthalpy



- The heat content a substance has at a given temperature and pressure
- Can’t be measured directly because there is no set starting point
- The reactants start with a heat content
- The products end up with a heat content
- So we can measure how much enthalpy changes

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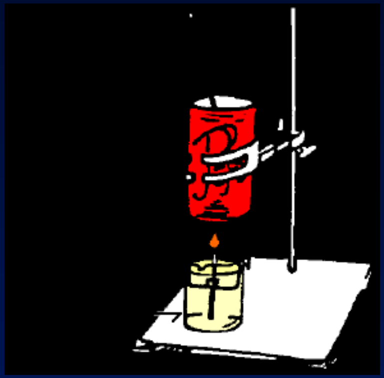
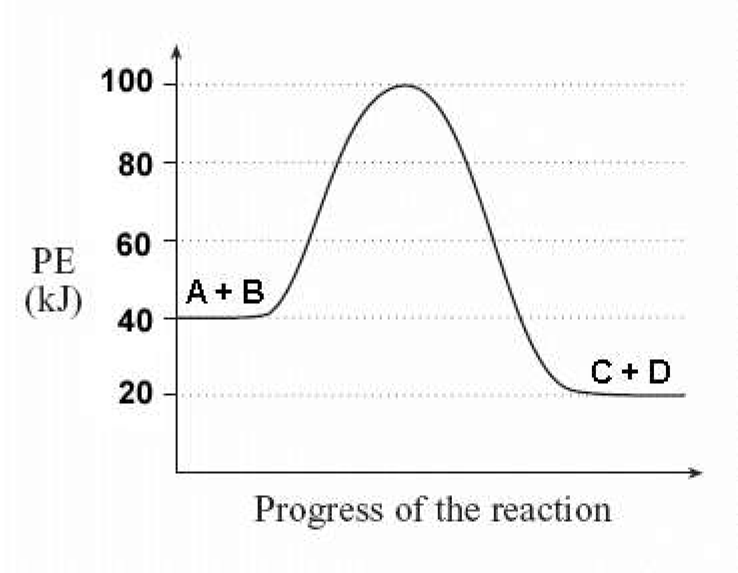
Enthalpy



- Symbol is H
- Change in enthalpy is ΔH
- Called, “delta H”
- If heat is released the heat content of the products is lower.
- ΔH is negative (exothermic)
- If heat is absorbed the heat content of the products is higher
- ΔH is positive (endothermic)

Exothermic

- The products are lower in energy than the reactants.
- Releases energy.
- ΔH is (-).



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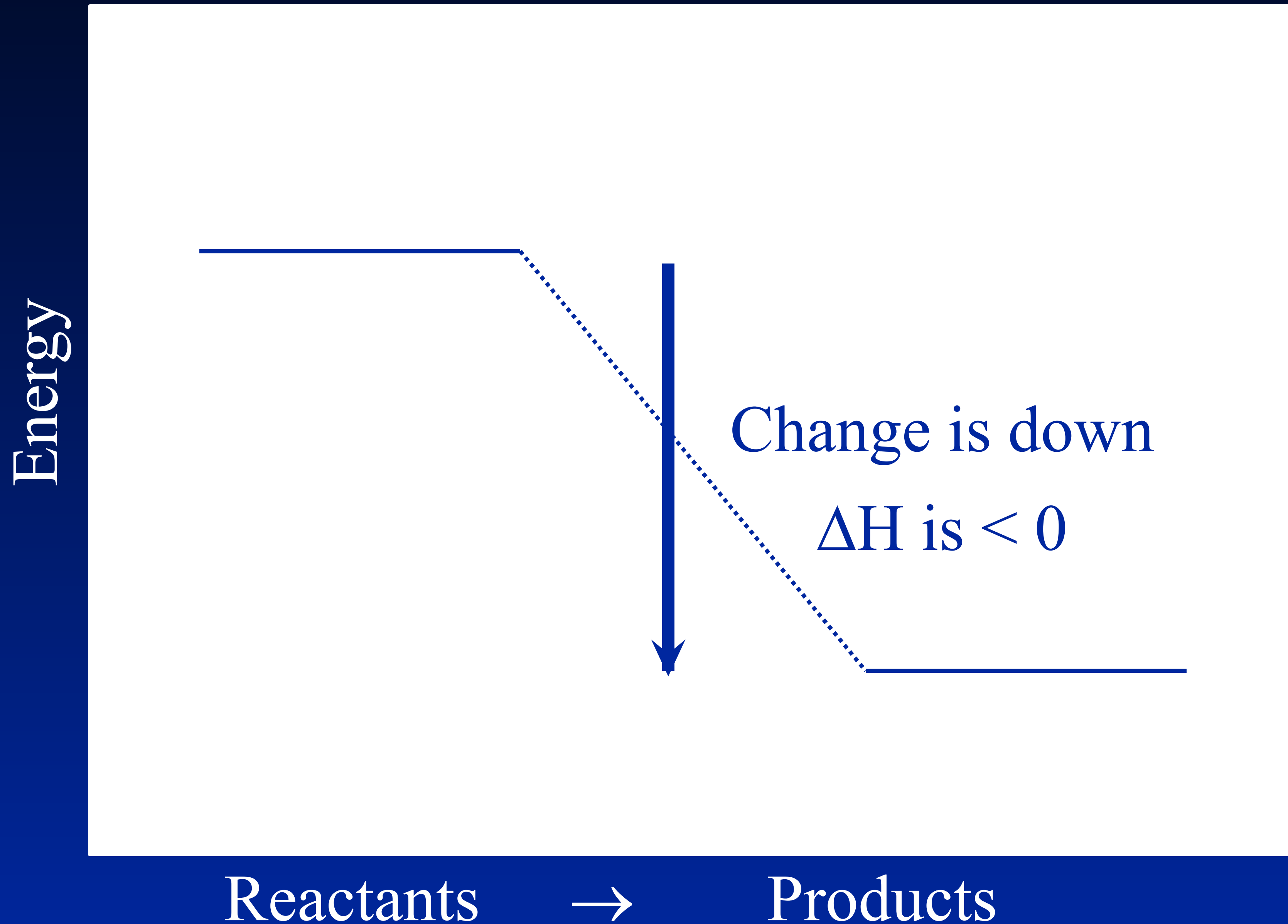

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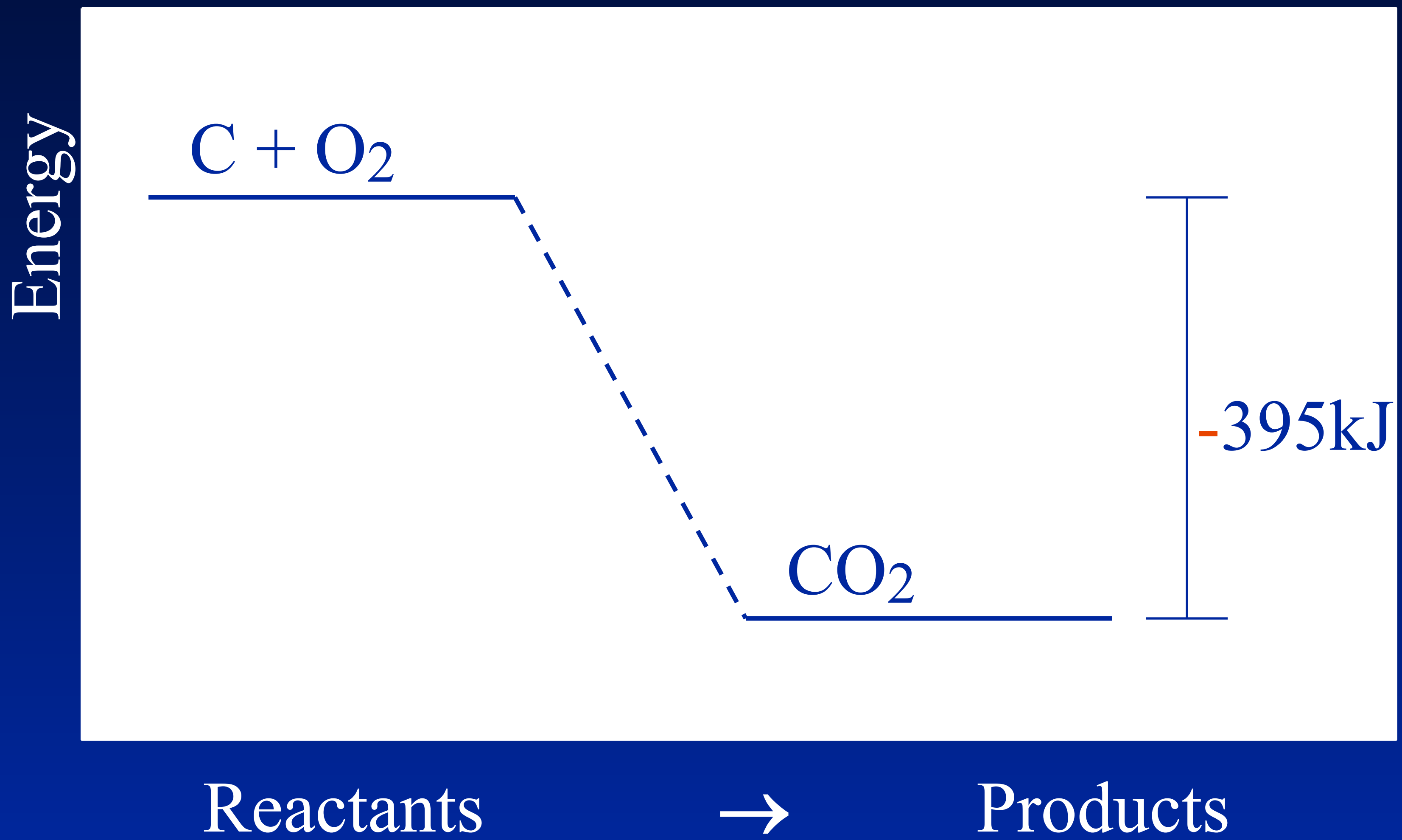
$$C + O_2 \rightarrow CO_2 + 395\ kJ$$



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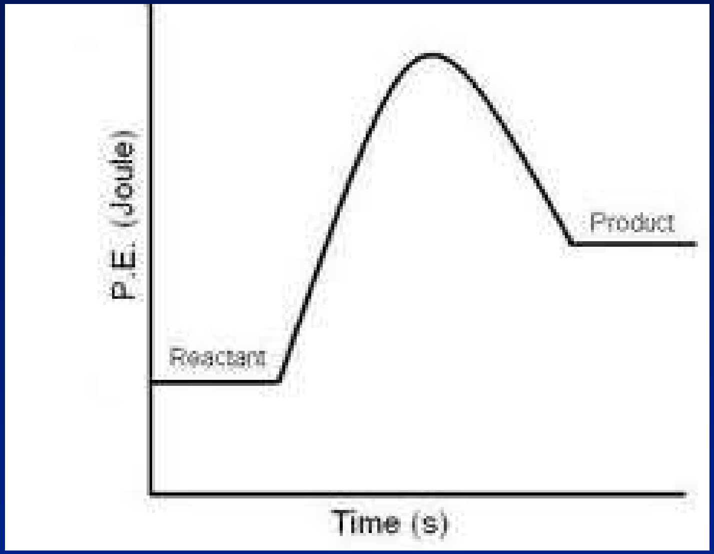
Endothermic

- The products are higher in energy than the reactants
- Absorbs energy



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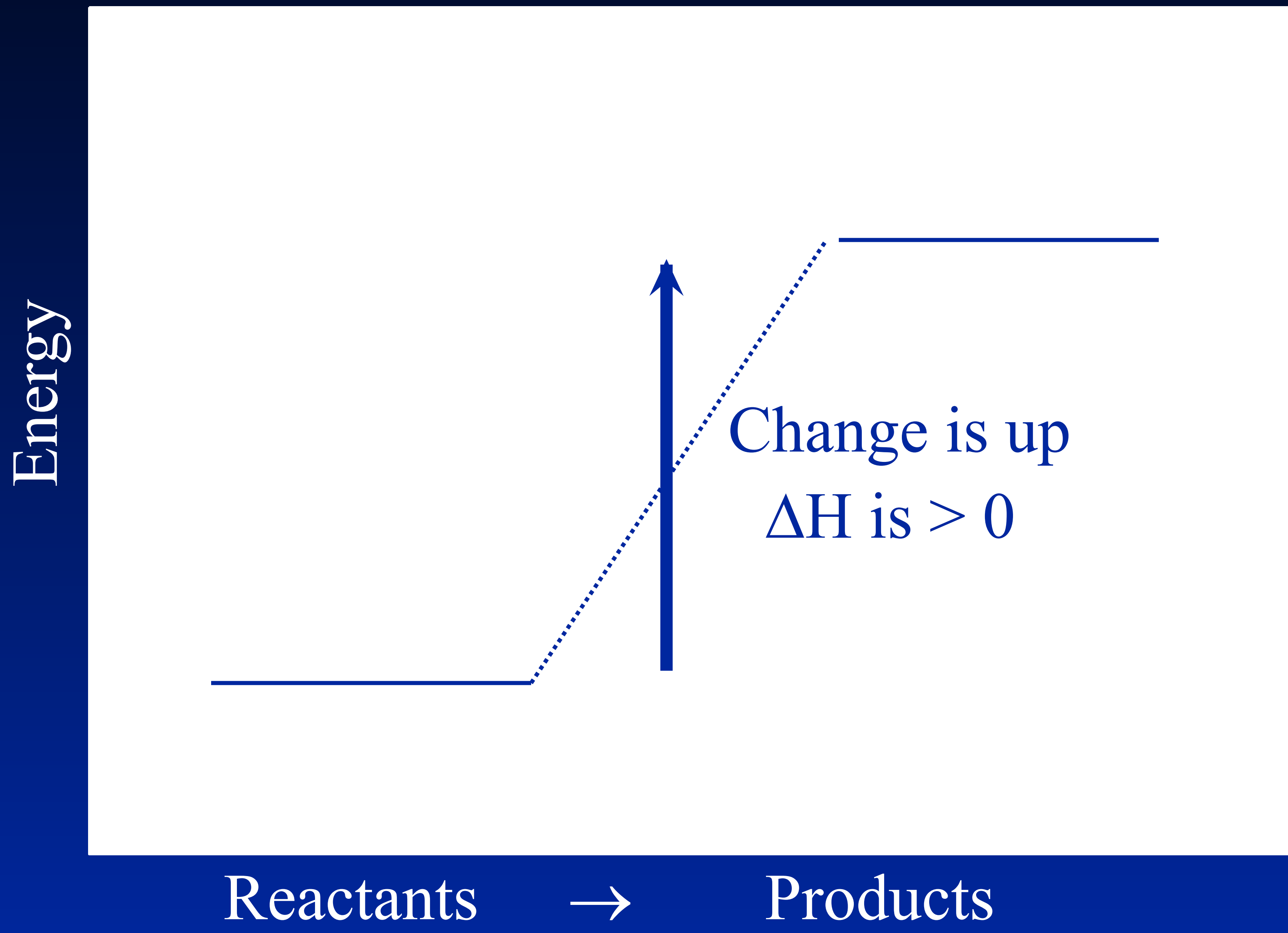

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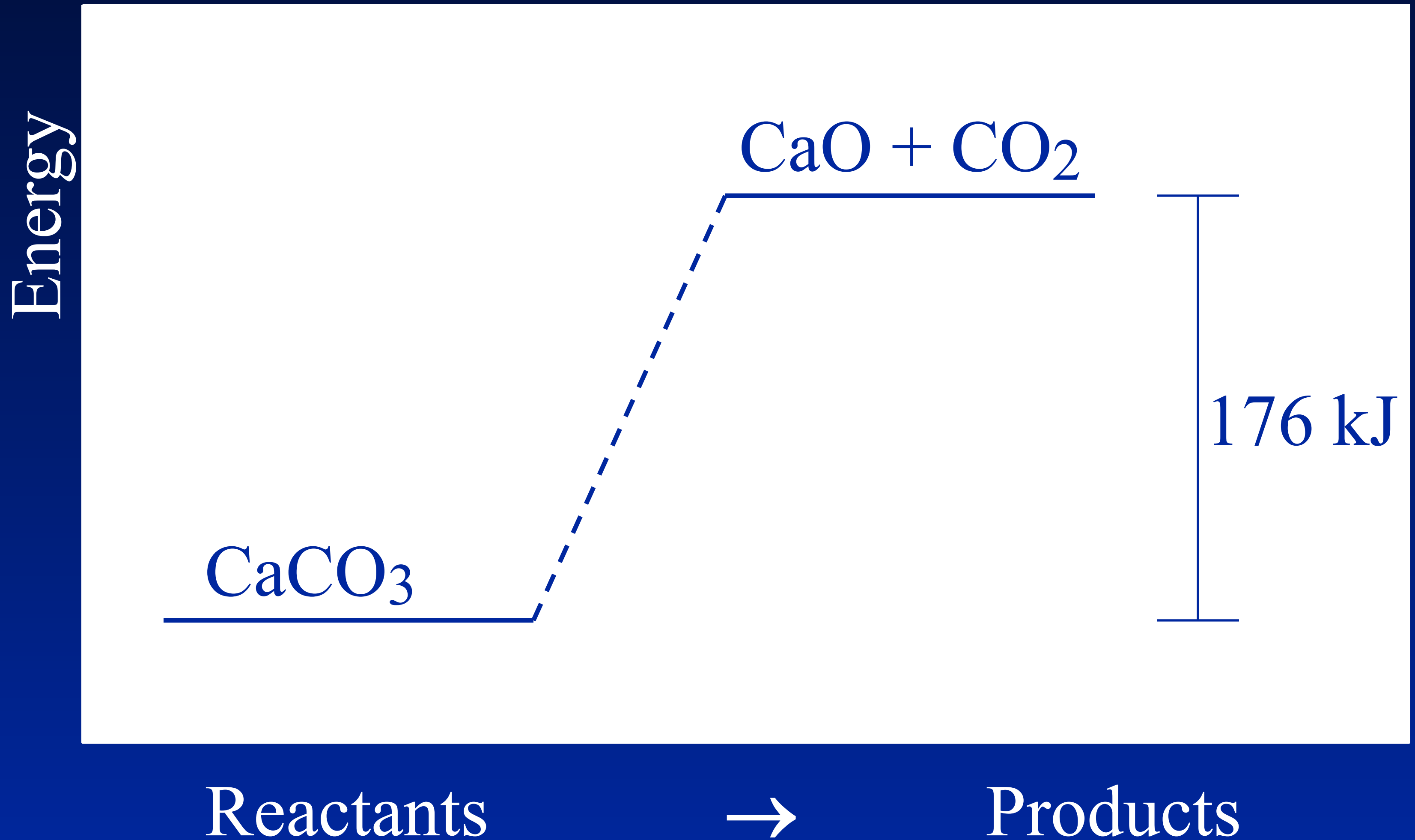
176 kJ + $CaCO_3 \rightarrow CaO + CO_2$



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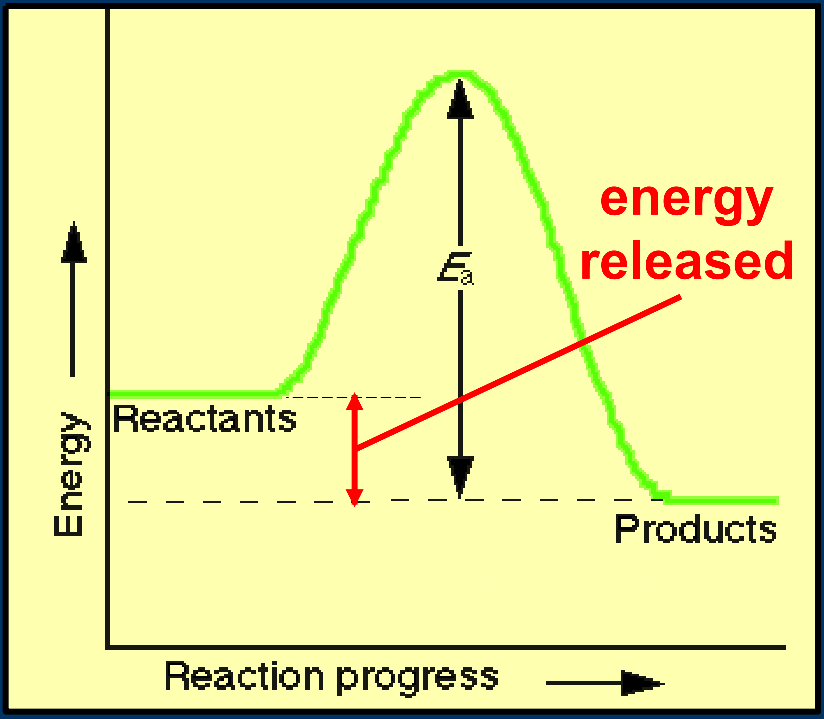
Exothermic Reaction

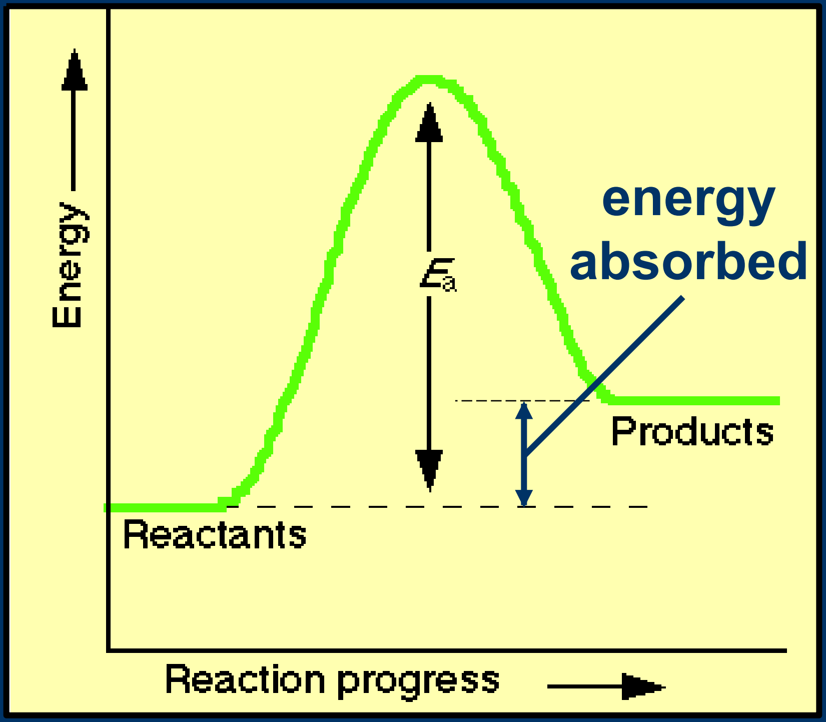
- reaction that releases energy
- products have lower PE than reactants



$$2H_2(l) + O_2(l) \rightarrow 2H_2O(g) + \text{energy}$$

Endothermic Reaction

- reaction that absorbs energy
- products have higher PE than products



$$2Al_2O_3 + \text{energy} \rightarrow 4Al + 3O_2$$

Heat of Reaction

- The heat that is released or absorbed in a chemical reaction.
- Equivalent to ΔH
- $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)} + 393.5$ kJ
- $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$ $\Delta H = -393.5$ kJ
- In a thermochemical equation it is important to say what the states are.
- $H_2(g) + 1/2O_2\,(g) \rightarrow H_2O(g)$ $\Delta H = -241.8$ kJ
  $H_2(g) + 1/2O_2\,(g) \rightarrow H_2O(l)$ $\Delta H = -285.8$ kJ

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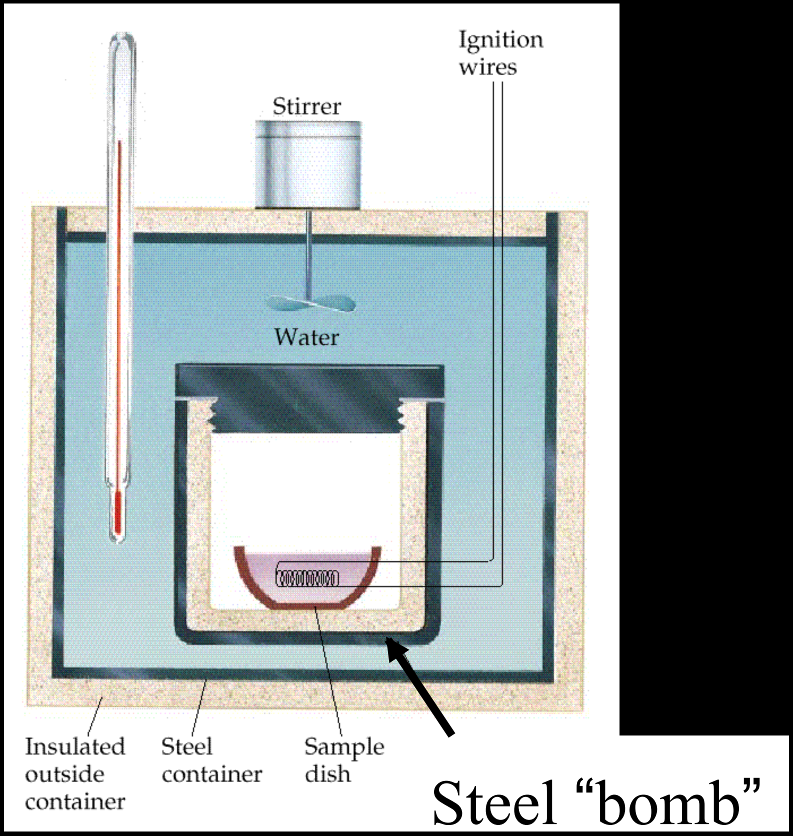
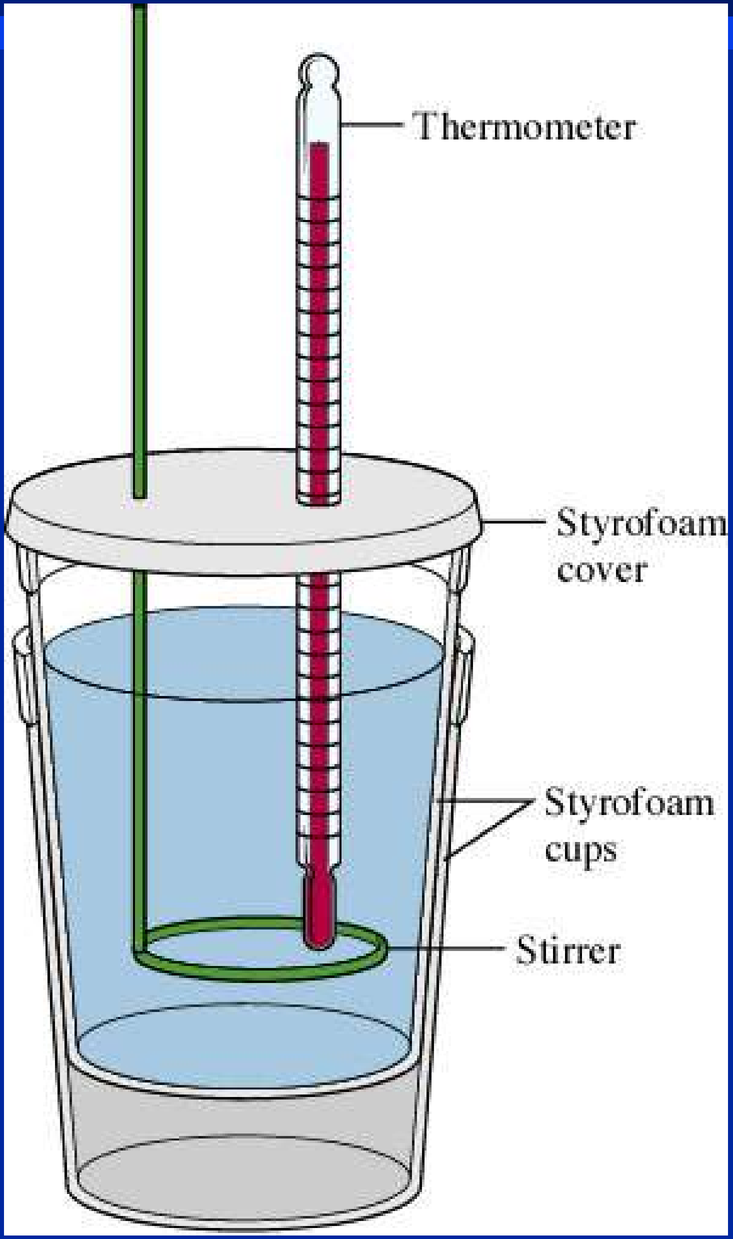

Ignition
wires
Stirrer
Water
Insulated
outside
container
Steel
container
Sample
dish
Steel "bomb"

A Coffee-Cup Calorimeter Made of Two Styrofoam Cups

We measure the energy content of something by change in temperature of the water (test question)



Entropy (randomness)

Entropy of a solid $<$ **Entropy of a liquid** $\ll$ **Entropy of a gas**

- A solid has an orderly arrangement.
- A liquid - molecules moving next to each other.
- A gas has molecules moving all over the place.

The End