

Chapter 5

Thermochemistry

The *energy* of chemical reactions

How do you keep track of it?

Where does it come from?

What is **Energy**?

- **The ability to:**

- do work
- Transfer heat

- **Work:**

- Energy that causes an object that has mass to move.

- **Heat:**

- Energy that causes the temperature of an object to rise.

Units of Energy

- The SI unit of energy is the joule (J).

$$1 \text{ J} = 1 \frac{\text{kg m}^2}{\text{s}^2}$$

- An older, non-SI unit is still in widespread use: The calorie (cal).

$$1 \text{ cal} = 4.184 \text{ J}$$

Energy has units of (mass)(velocity)²

Remember **kinetic energy** = $1/2mv^2$

Units of Energy

- The SI unit of energy is the joule (J).

$$1 \text{ J} = 1 \frac{\text{kg m}^2}{\text{s}^2}$$

Energy has units of (mass)(velocity)²

Which of the following represent units of E?

- A. lb in/h²
- B. pint ft/sec²
- C. ton mile²/year²
- D. None of the above
- E. A and C

Work

- Energy used to move an object over some distance.
- $w = F \cdot d$,
 $w = \text{work}$,
 $F = \text{force}$
 $d = \text{distance over which force exerted.}$

Note units:

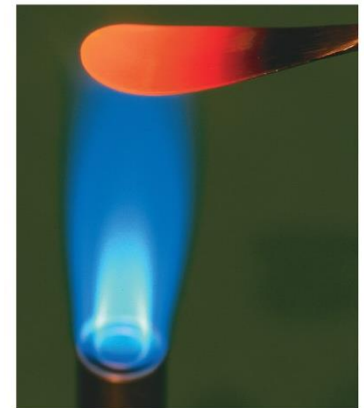
$$F = ma, \text{ mass}(\text{distance}/\text{sec}^2)$$

$$W = F(d) = \text{mass}(\text{distance}^2/\text{sec}^2)$$

$$= mv^2$$



(a)



(b)

Kinetic Energy

Energy an object possesses by virtue of its motion.

$$KE = \frac{1}{2} mv^2$$



Heat

- Energy can also be transferred as **heat**.
- **Heat** flows from **warmer (higher T)** stuff to **cooler (lower T)** stuff.
- T is proportional to Average Kinetic Energy.



Potential Energy

Energy an object possesses by virtue of its position or chemical composition.

Energy that could *potentially be transferred*.

More potential E

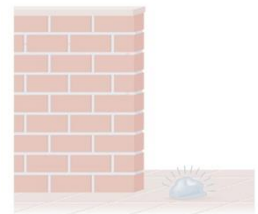
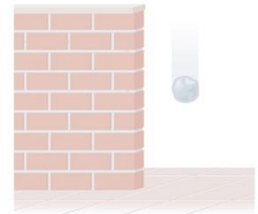
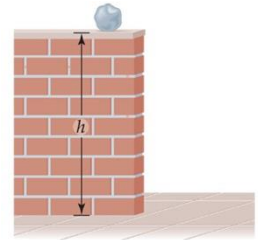


Less P.E. as bike goes down.



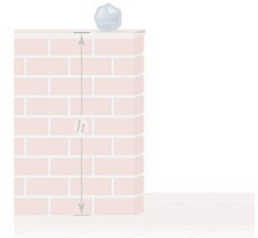
Transfer of Energy

- a) Add P.E. to a ball by lifting it to the top of the wall



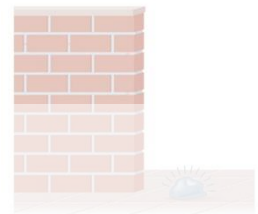
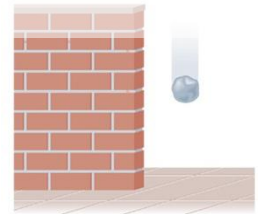
Transferal of Energy

a) Add P.E. to a ball by lifting it to the top of the wall



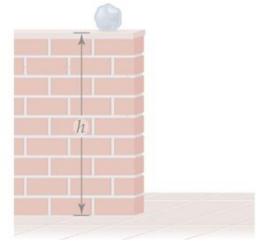
b) As the ball falls,

P.E -----> K. E. ($\frac{1}{2}mv^2$)



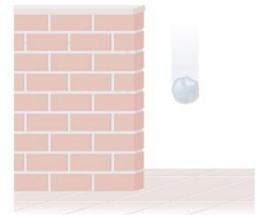
Transfer of Energy

a) Add P.E. to a ball by lifting it to the top of the wall



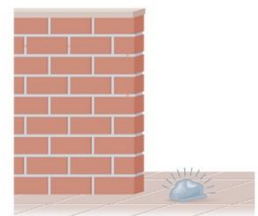
b) As the ball falls,

P.E. -----> K. E. ($\frac{1}{2}mv^2$)



Ball hits ground, K.E. =0, but E has to go somewhere. So

1. Ball gets squashed
2. Heat comes out.
3. Ground also releases heat



Energy accounting

- We must identify where different types of energy go.
- Therefore, we must identify the *places*.

System and Surroundings



- The **system** includes the stuff we want to study (here, the hydrogen and oxygen molecules).
- The **surroundings** are everything else (here, the cylinder, piston and everything else).

Internal Energy

The internal energy of a system is the sum of all kinetic and potential energies of all components of the system; we call it E .

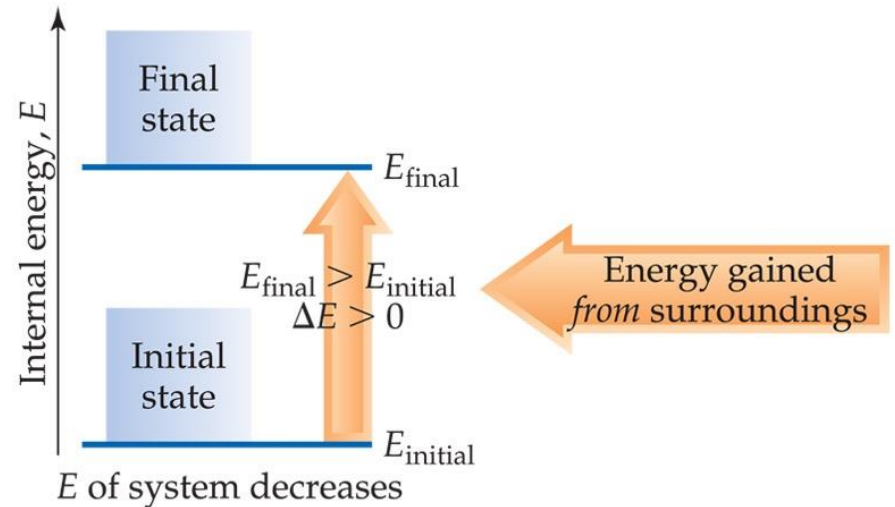
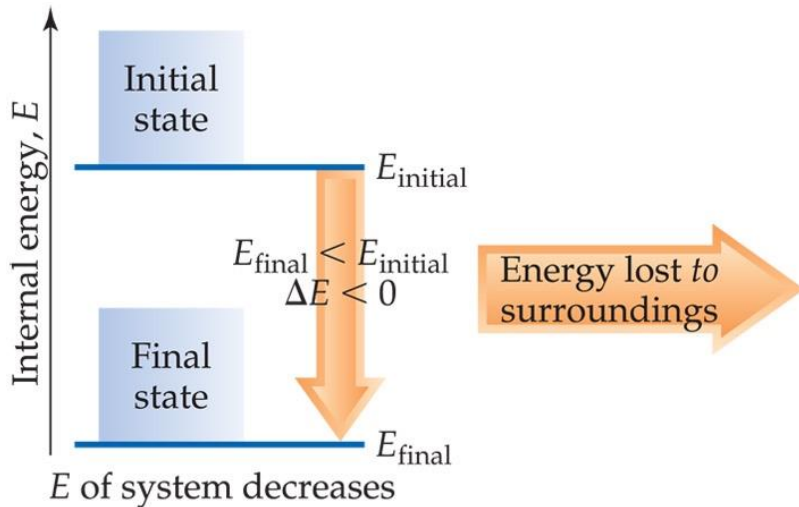
$$E_{\text{internal,total}} = E_{\text{KE}} + E_{\text{PE}} + E_{\text{electrons}} + E_{\text{nuclei}} + \dots$$

Almost impossible to calculate total internal energy
Instead we always look at the **change** in energy (ΔE).

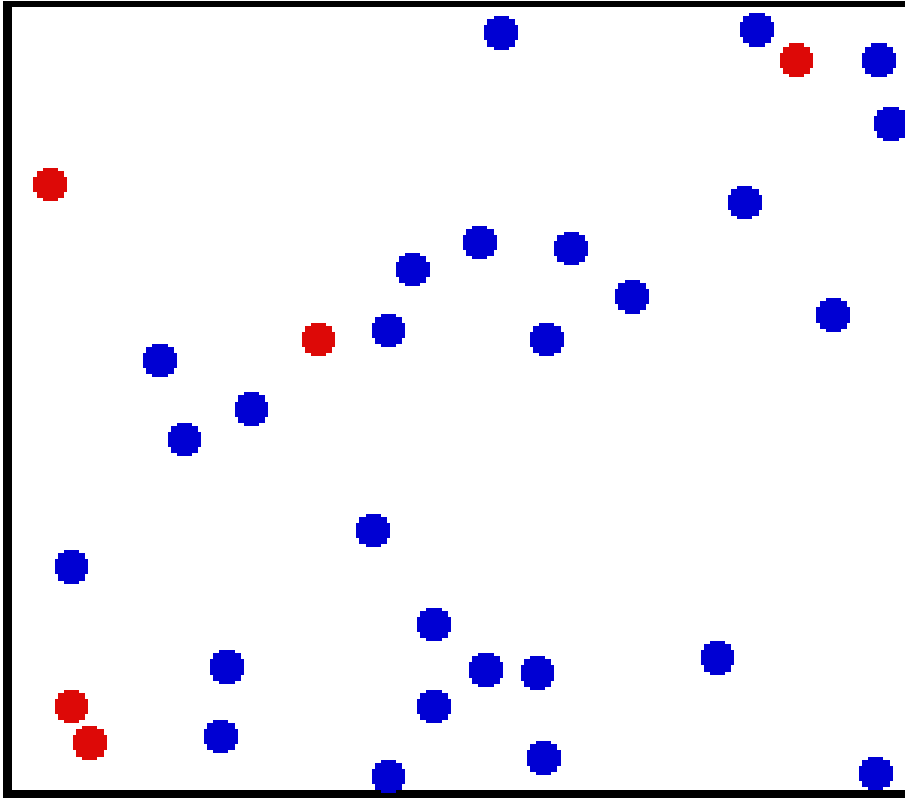
Internal Energy

By definition, the change in internal energy, ΔE , is the final energy of the system minus the initial energy of the system:

$$\Delta E = E_{\text{final}} - E_{\text{initial}}$$

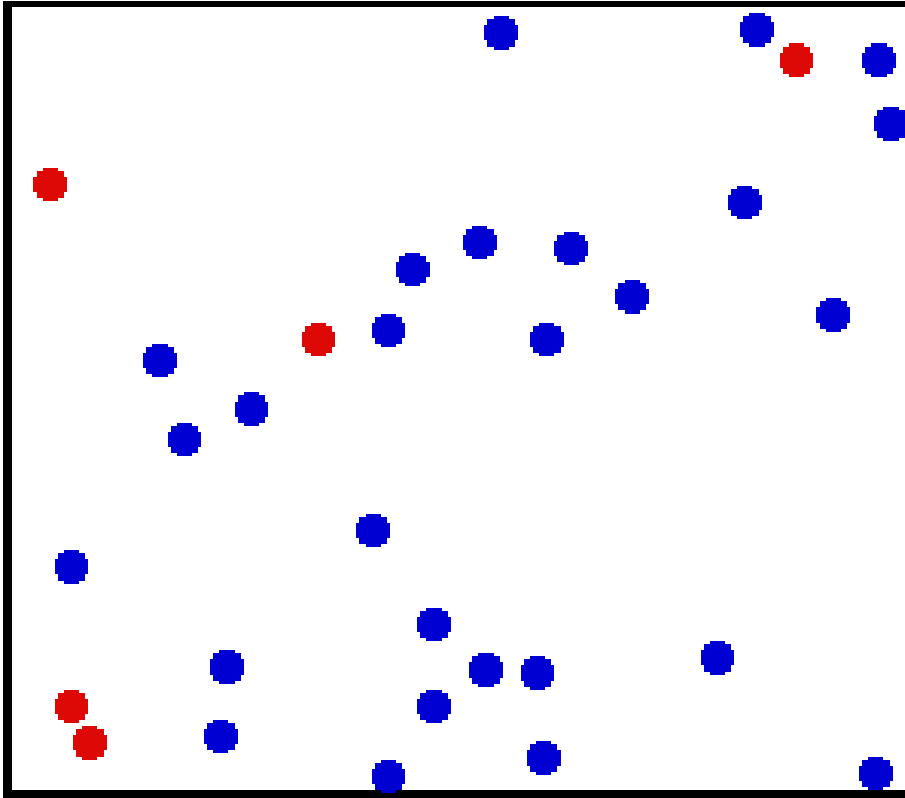


Heat review



- Heat is transfer of *random* kinetic energy from:
- Stuff with **higher K.E. (higher T)** to stuff with **lower K.E.**
- When faster atoms/molecules hit slower ones:
- The faster ones get slower and the slower get faster.

Heat.



- Transfer of *random* atomic/molecular motion

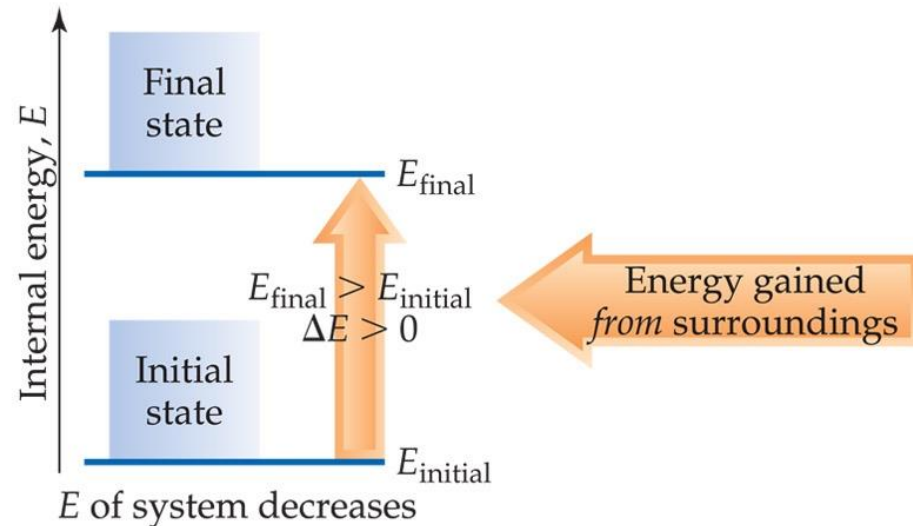
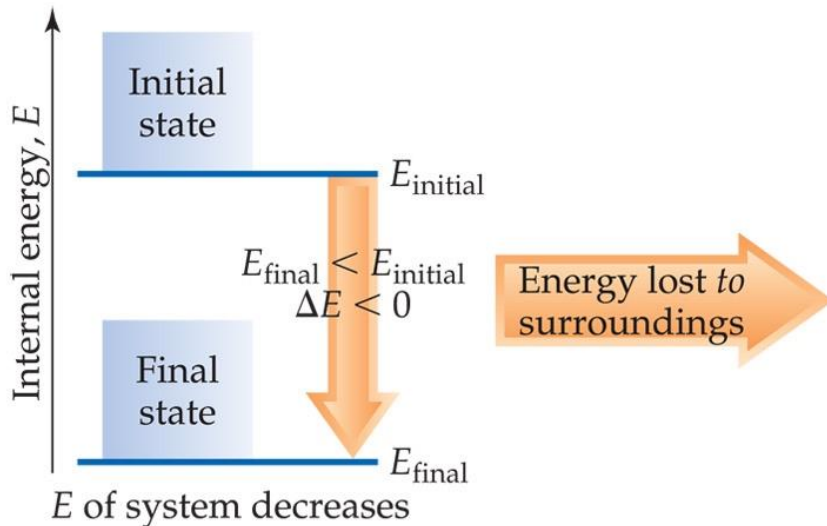
work



- Transfer of *organized* atomic/molecular motion.

First Law of Thermodynamics

- Energy is **conserved**.
- In other words, the total energy of the universe is a constant; $\Delta E_{\text{System}} = -\Delta E_{\text{surroundings}}$



Changes in Internal Energy

- If $\Delta E > 0$, $E_{\text{final}} > E_{\text{initial}}$

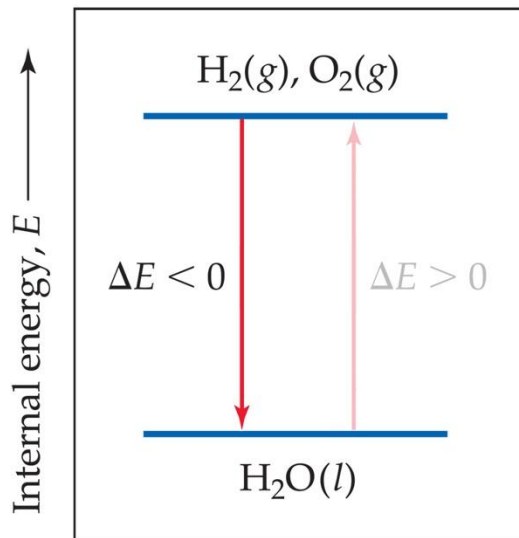
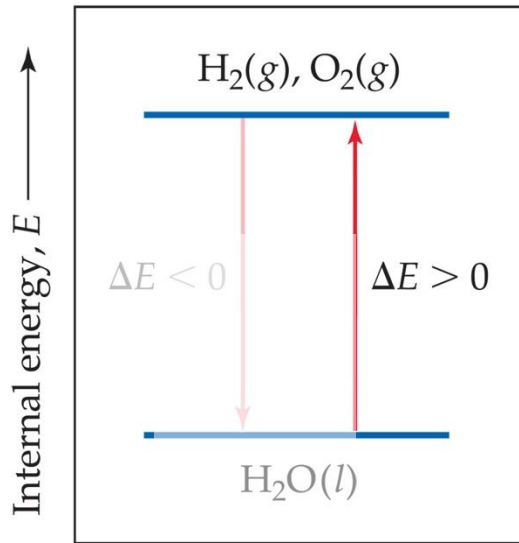
➤ Therefore, the system *absorbed* energy from the surroundings.

➤ This energy change is called **endergonic**.

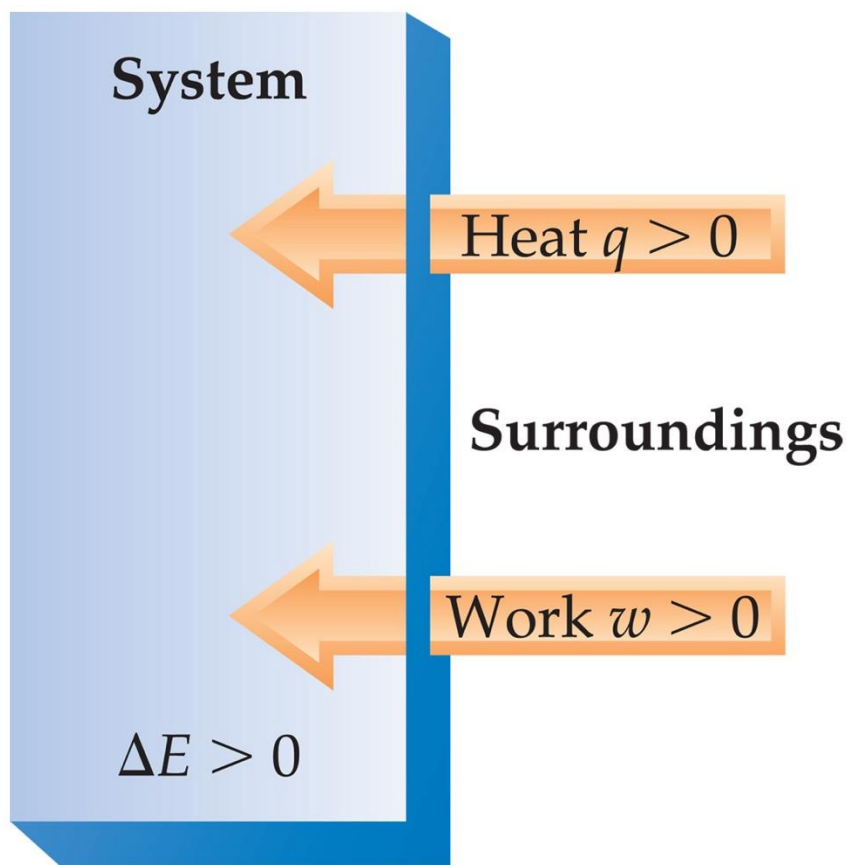
- If $\Delta E < 0$, $E_{\text{final}} < E_{\text{initial}}$

➤ Therefore, the system *released* energy to the surroundings.

➤ This energy change is called **exergonic**.



Changes in Internal Energy



- When energy is exchanged between the system and the surroundings, it is exchanged as either heat (q) or work (w).
- That is,

$$\square \Delta E = q + w$$

ΔE , q , w , and Their Signs

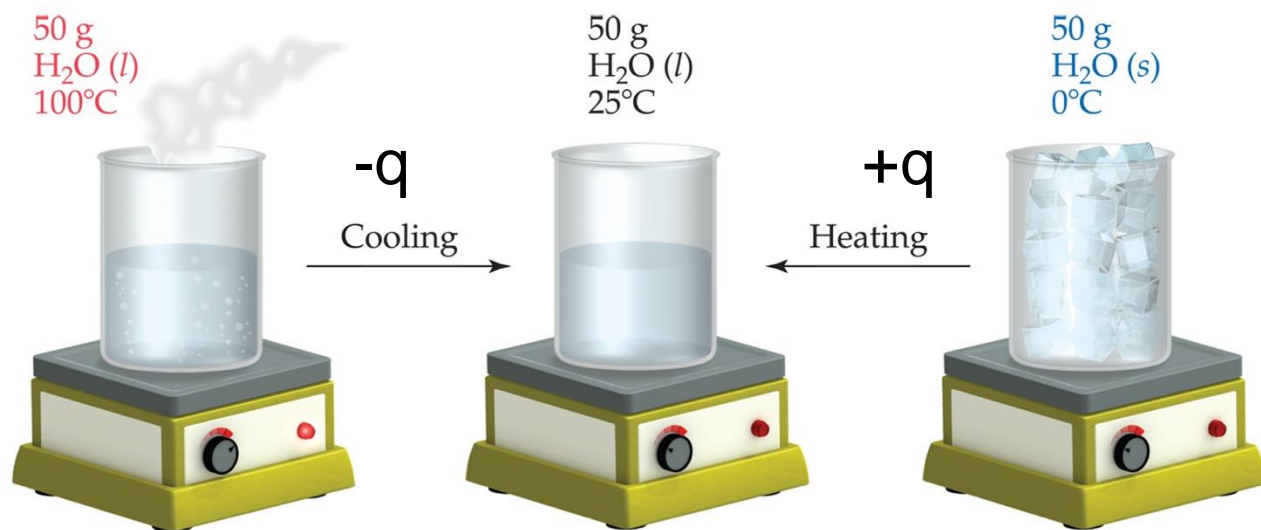
For q + means system *gains* heat

- means system *loses* heat

For w + means work done *on* system

- means work done *by* system

For ΔE + means *net gain* of energy by system - means *net loss* of energy by system



Surroundings
suck heat out of
water.

hot plate adds
heat to water

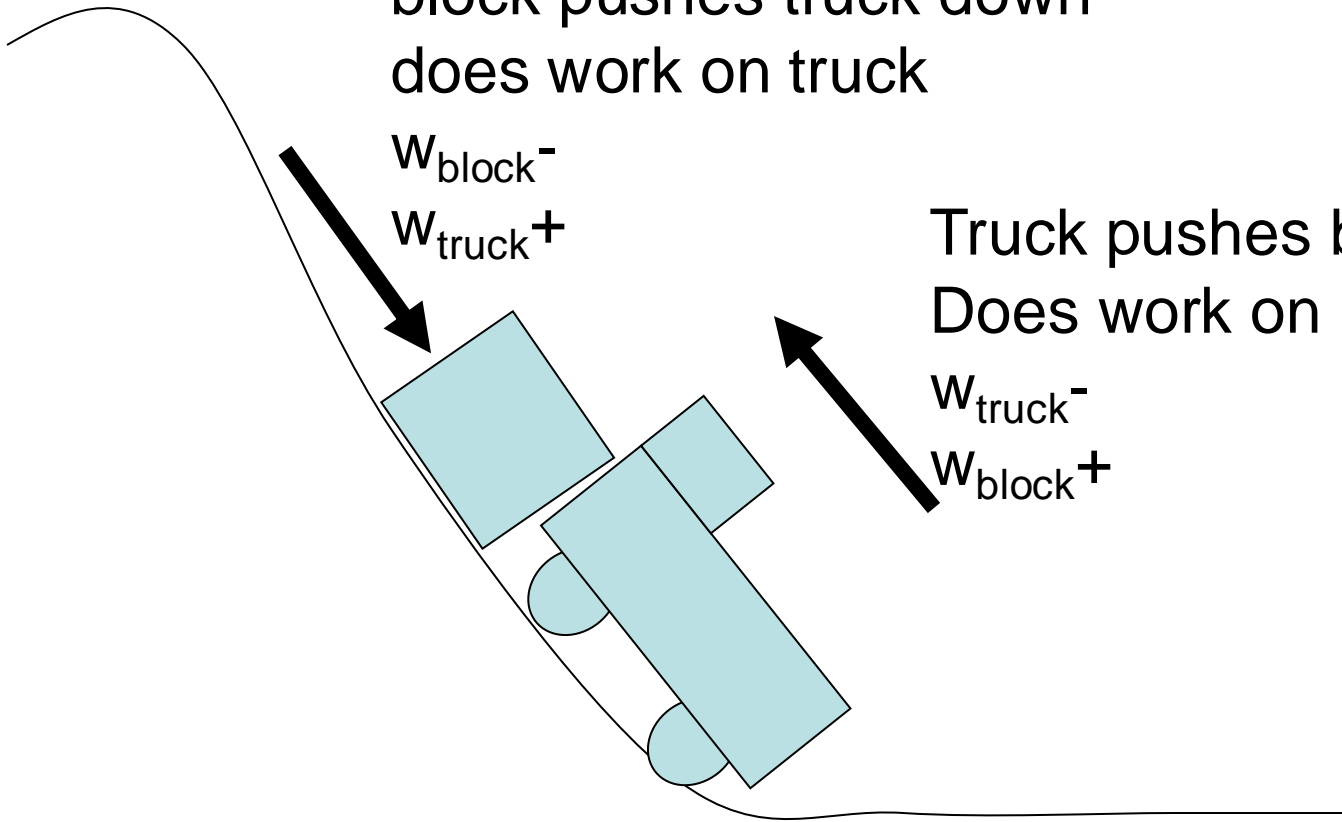
Sign of work

block pushes truck down
does work on truck

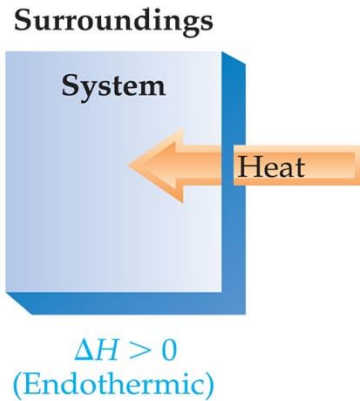
$$W_{\text{block}}^-$$
$$W_{\text{truck}}^+$$

Truck pushes block up.
Does work on block

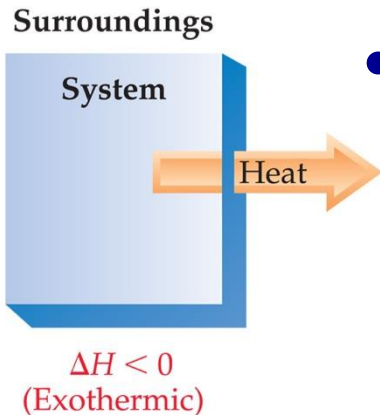
$$W_{\text{truck}}^-$$
$$W_{\text{block}}^+$$



Exchange of Heat between System and Surroundings



Heat absorbed by system from surroundings, is ***endothermic***.



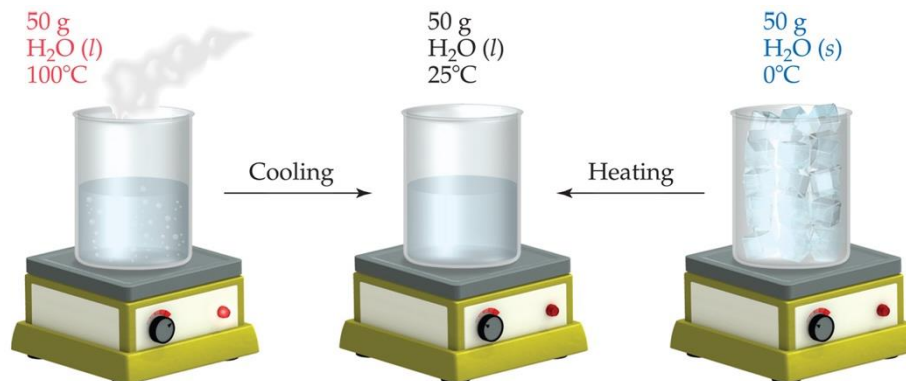
- Heat released by system to surroundings, is ***exothermic***.

State Functions

Total internal energy of a system:

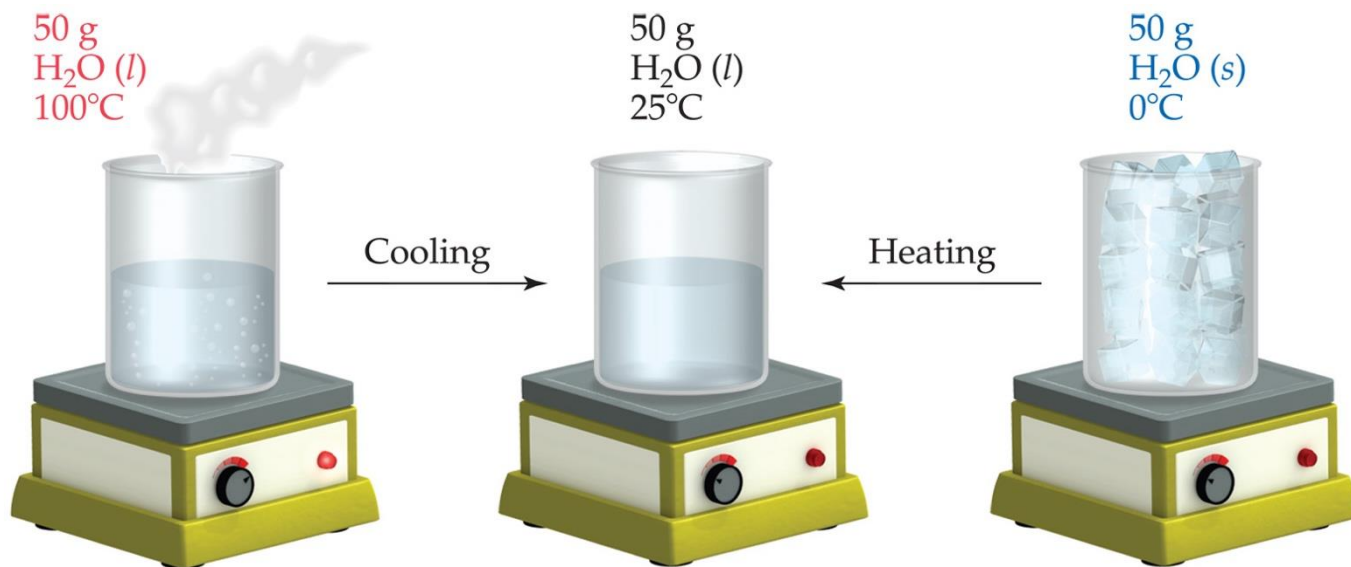
$$\text{K.E.} + E_{\text{electrons}} + E_{\text{nucleus}} + \text{P.E.}_{\text{total}}$$

virtually impossible to measure/calculate



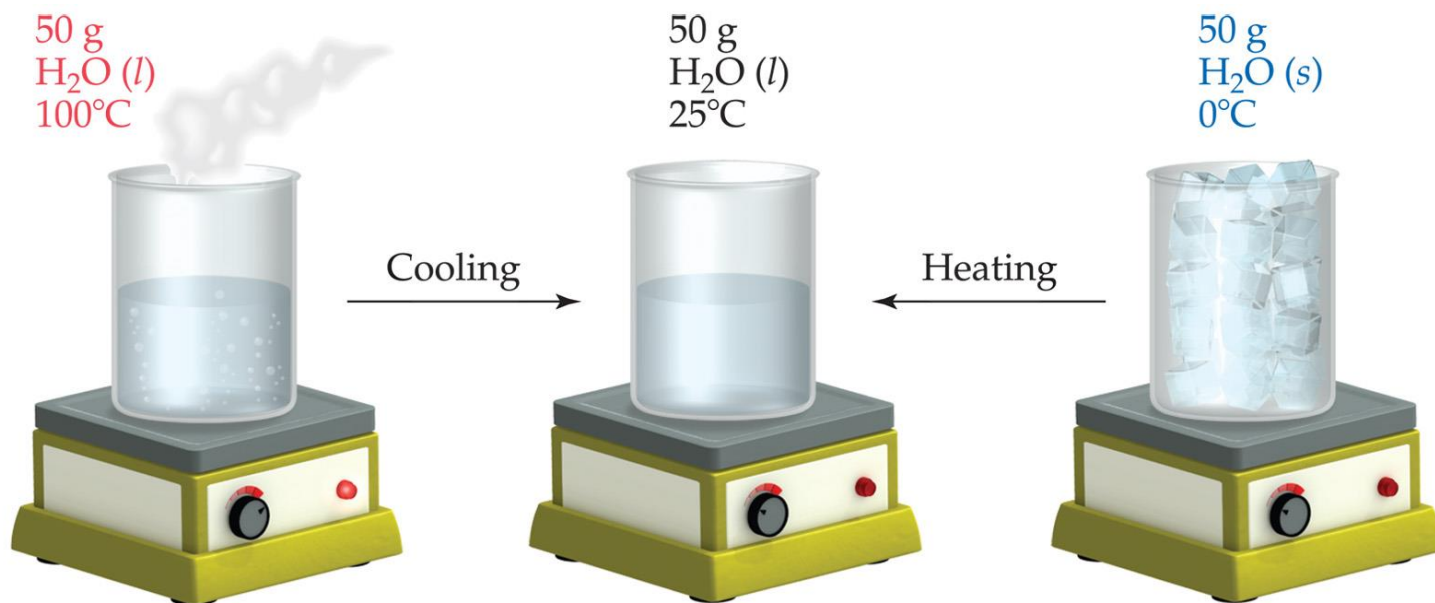
State Functions

- However, the internal energy of a system is **independent** of the *path* by which the system achieved that state.
 - In the system below, the water could have reached room temperature from either direction.



State Functions

- Therefore, internal energy is a ***state function***.
- because its **PATH INDEPENDENT**
- And so, ΔE depends only on E_{initial} and E_{final} .



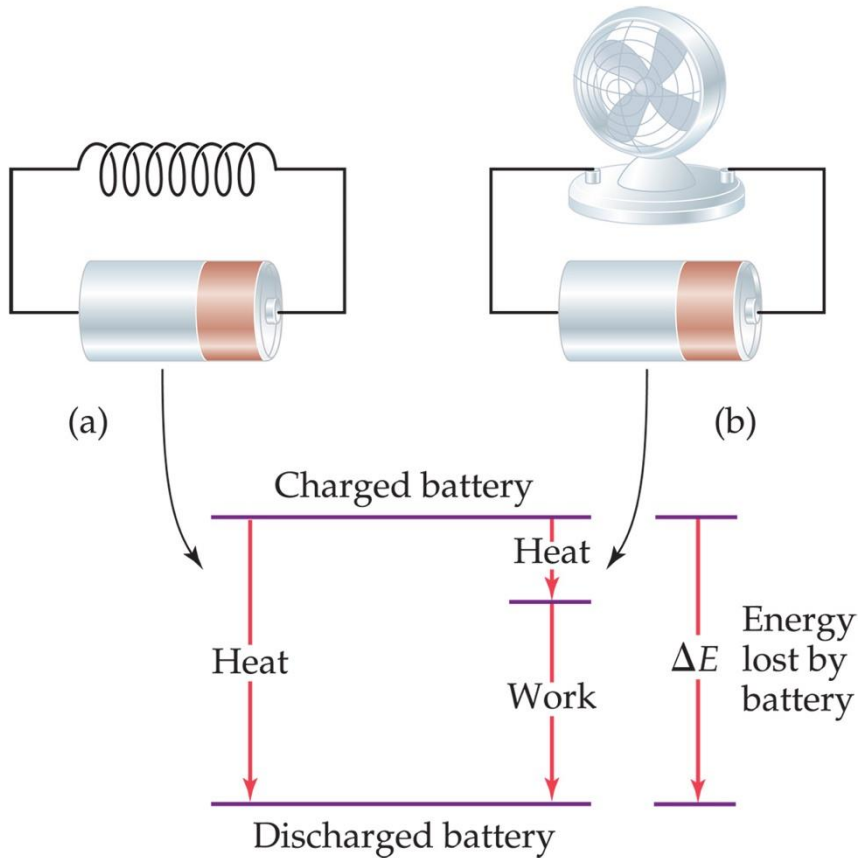
State Functions

- Which of the following are state functions?
 - A. Temperature
 - B. Volume
 - C. mass
 - D. All of the above
 - E. A and B

State Functions

All heat.

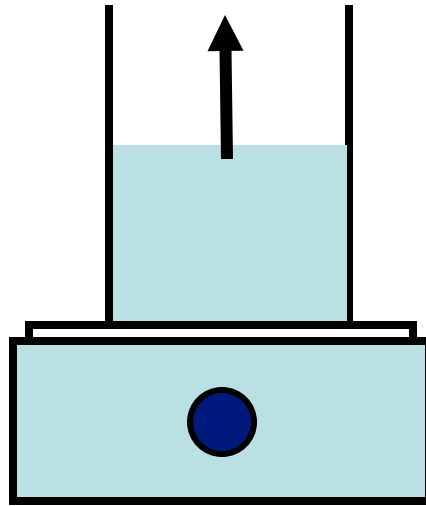
Work and heat



- However, **q** and **w** are **not** state functions.
- Whether the battery is connected to a heater or is discharged by running the fan, its ΔE is the same.
 - But q and w are different in the two cases.

Work, the chemistry kind.

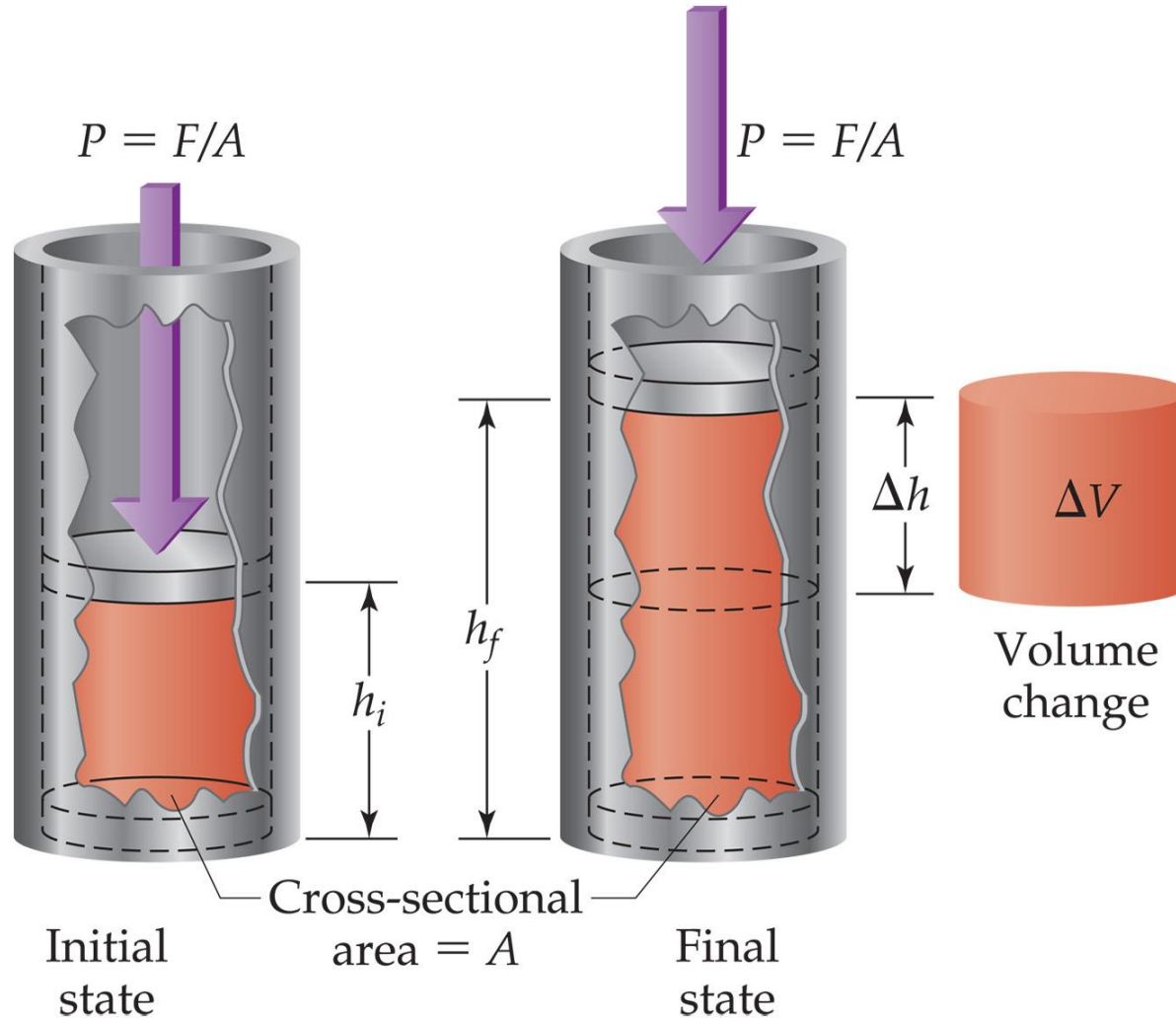
How does chemistry do work?



process in an open container (chemical reaction in a beaker)
w? (can there be any work)?

Yes, **evolving gases** could push on the surroundings.

Catch the work, do the same process in a cylinder



Process evolves gas, pushes on piston, work done on piston.

The internal combustion engine

Catch the work, do the same process in a cylinder

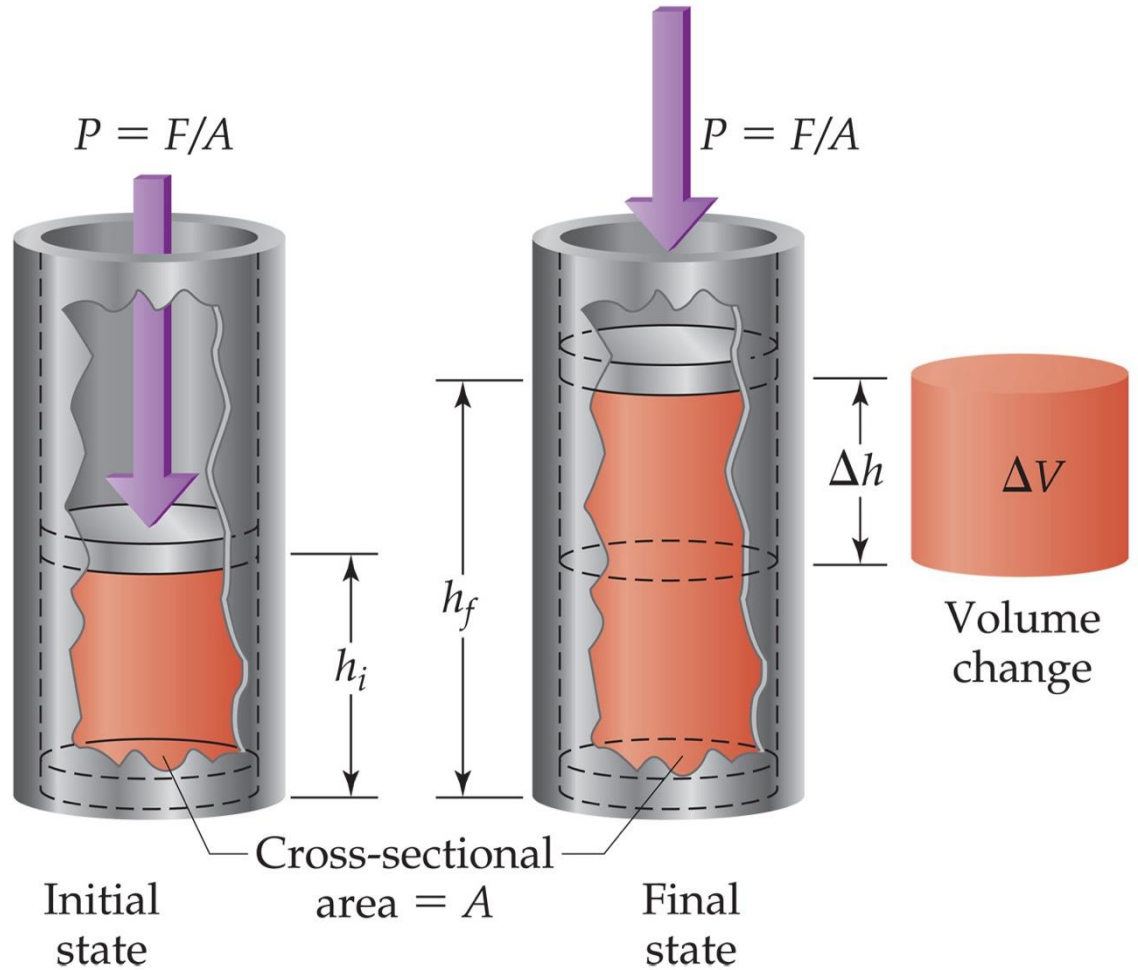
$$w = F \cdot d, \quad F = P \cdot A, \quad d = \Delta h$$

$$w = -P \cdot A \Delta h = -P \Delta V$$

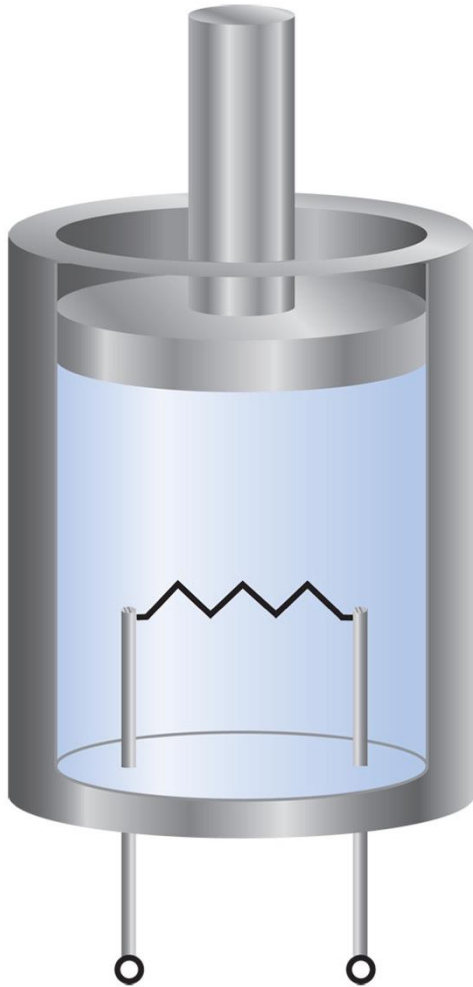
Negative because an **increase** in Volume means that the system is doing work **on** the surroundings.

$$\Delta E = q + w = q - P \Delta V$$

$$q_p = \Delta E + P \Delta V$$

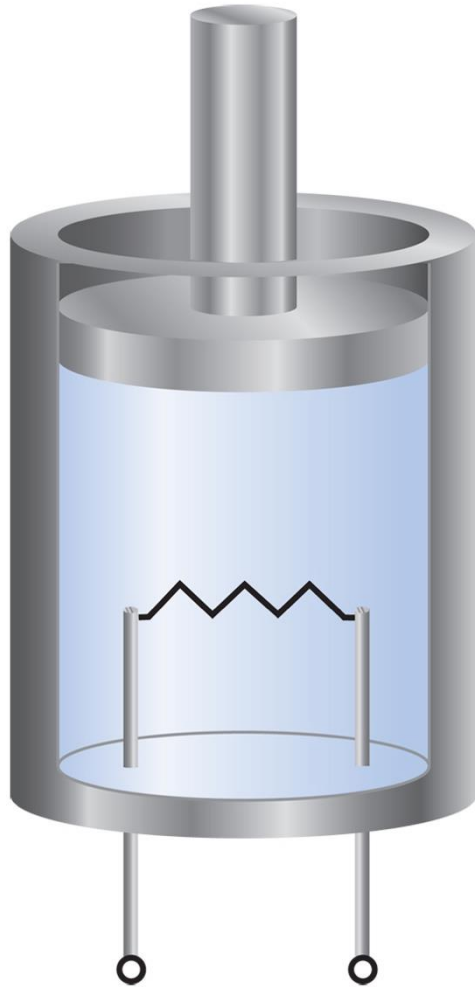


Example



- Gas inside cylinder with electric heater. Add 100 J heat with heater. Two scenarios:
 - 1. Piston can go up and down
 - 2. Piston stuck.
- a. What happens to T in each case?
- b. What about q and w for each case?
- c. What about ΔE in each case?

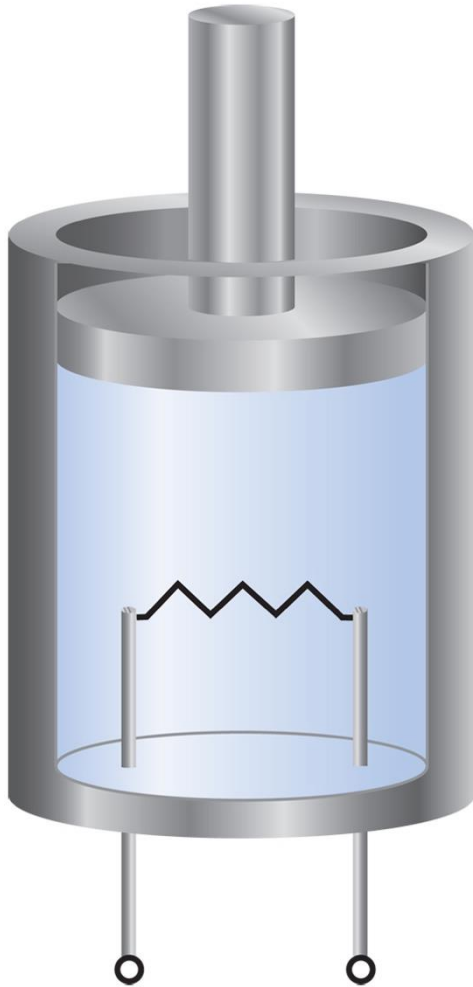
Example



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- Gas inside cylinder with electric heater. Add 100 J heat with heater. Two scenarios:
 - 1. Piston can go up and down
 - 2. Piston stuck.
- a. **What happens to T in each case?**
 - A. T goes up the same
 - B. T goes down the same
 - C. T goes up more in 1
 - D. T goes up more in 2
 - E. T goes down more in 2

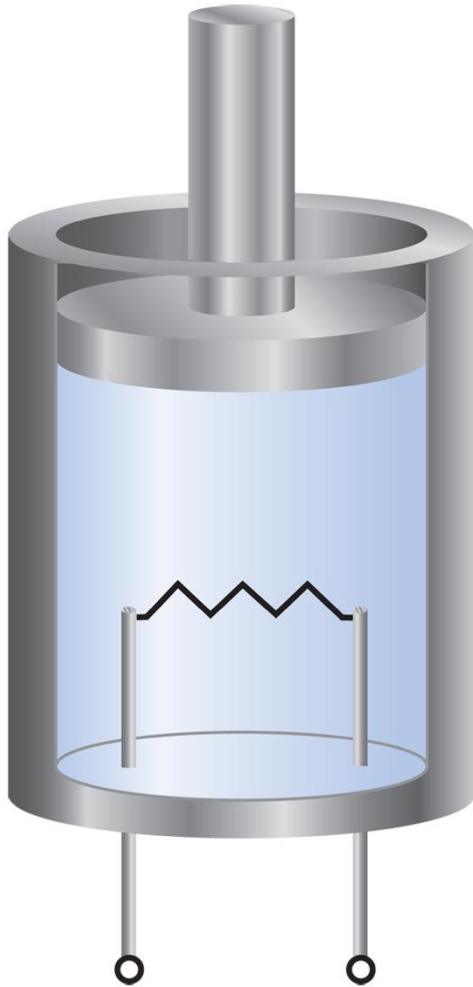
Example



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- Gas inside cylinder with electric heater. Add 100 J heat with heater. Two scenarios:
 - 1. Piston can go up and down
 - 2. Piston stuck.
- b. What about q for each case?
 - A. q the same in 1 and 2
 - B. More q in 1 than 2
 - C. More q in 2 than 1
- c. What about w for each case?
 - A. w the same in 1 and 2
 - B. More w done in 1 than 2
 - C. More w done in 2 than 1

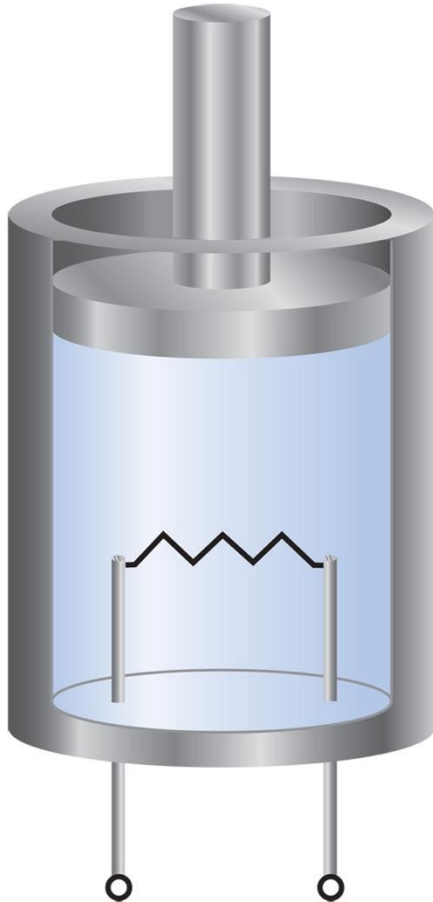
Example



- Gas inside cylinder with electric heater. Add 100 J heat with heater. Two scenarios:
 - 1. Piston can go up and down
 - 2. Piston stuck.
- d. What about ΔE in each case?
 - A. Same in each case
 - B. More in 1 than 2
 - C. More in 2 than 1

Example

- Gas inside cylinder with electric heater.
- Add 100 J heat with heater.
- 1. Piston can go up and down
- 2. Piston stuck.
- a. What happens to T in each case?
- b. What about q and w for each case?
- c. What about ΔE in each case?

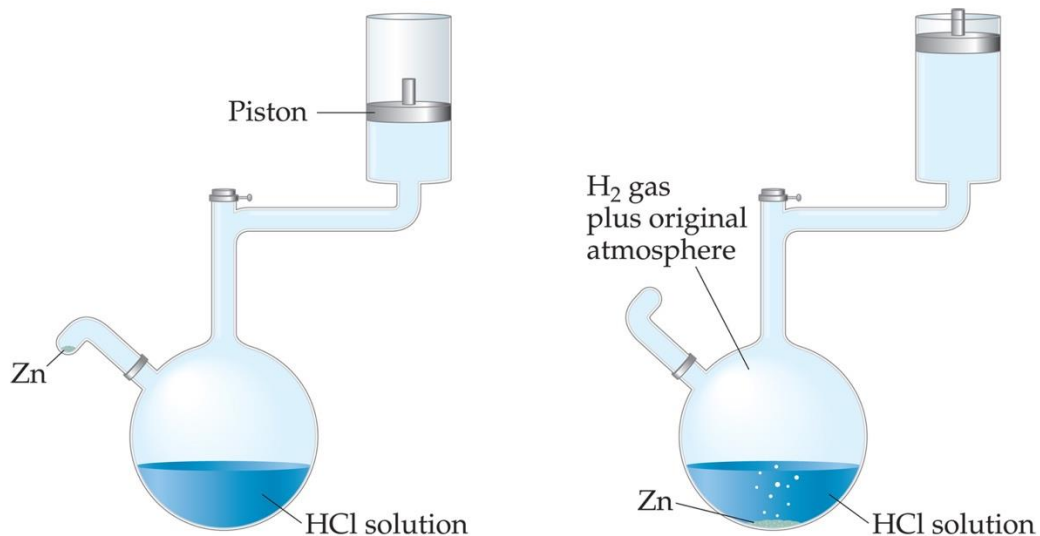


- a.1. Piston goes up, some E goes to expand gas, do work. T goes up less
- a.2 T goes up more, all E goes to q .
- b.1. q positive w negative if piston goes up.
- b.2. $w = 0$, q larger
- c. ΔE the same & + in each case

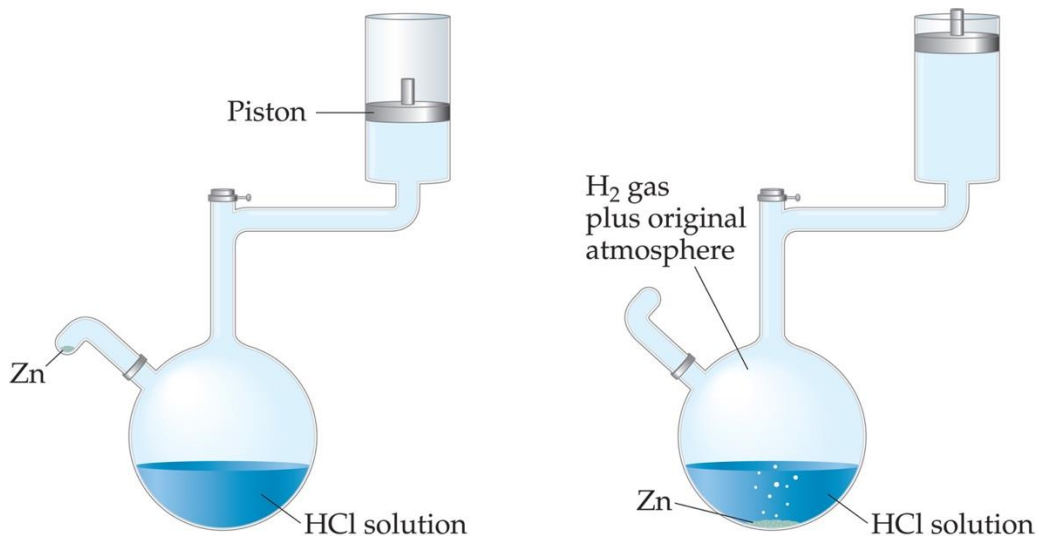
Work

Now we can measure the work:

$$w = -P\Delta V$$

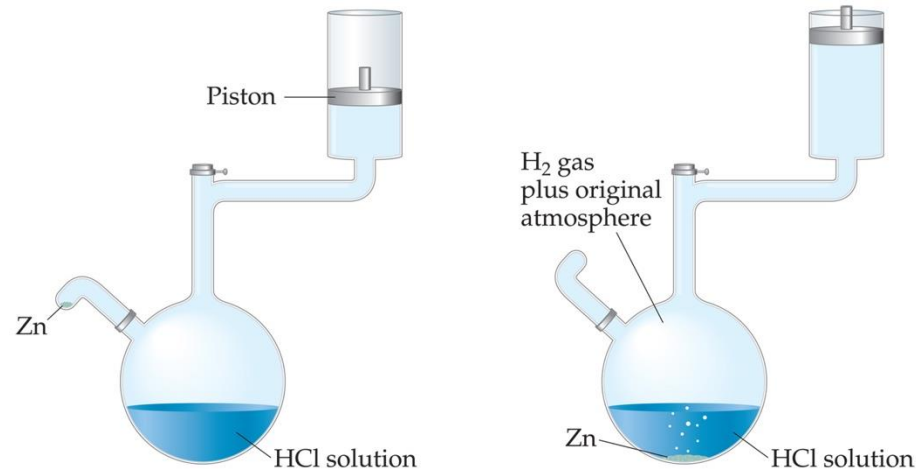


Work



1 mole of Zn reacts. How much work is done ($P = 1 \text{ atm}$, density of $\text{H}_2 = 0.0823 \text{ g/L}$)?

Work



1 mole of Zn reacts. How much work is done ($P = 1 \text{ atm}$, density of $\text{H}_2 = 0.0823 \text{ g/L}$)?
1 mole of $\text{H}_2(\text{g})$ is produced.



1mol

1 mol

2.014 g/mol

2.014 g

$d=m/V$

$V=m/d$

$$V = 2.014\text{g}/0.0823\text{g/L} = \mathbf{24.47 \text{ L}}$$

$$W = -P\Delta V = -1\text{atm}(\mathbf{24.47\text{L}}) = -24.47 \text{ L(atm)}. \quad \text{L(atm)} = \text{dm}^3(101,325 \text{ kgm}^{-1}\text{sec}^{-2}) \\ = 101325 \text{ kg}(10^{-1})\text{dm}^2\text{sec}^{-2} \text{ has units of } \text{mv}^2.$$

Enthalpy(H)

$$H = E + PV$$

**This is the definition of Enthalpy for *any* process
Buy why do we care?**

Enthalpy(H)

$$H = E + PV$$

- at constant pressure, ΔH , is
(Δ = change in thermodynamics)

$$\Delta H = \Delta(E + PV) = \Delta E + \Delta PV$$

- If P constant:

$$\Delta H = \Delta E + P\Delta V$$

Enthalpy

- Since $\Delta E = q + w$ and $w_p = -P\Delta V$ (P const.) substitute these into the enthalpy expression:

$$\Delta H_p = \Delta E_p + P\Delta V$$

$$\Delta H_p = (q_p + w_p) - w_p$$

$$\Delta H_p = q_p$$

- **Note: true at constant pressure**
- ***q is a state function* at const P & only PV work.**

$$H = E + PV$$

- **Because:**
- If pressure is constant (like open to atmosphere, i.e. most things) and

$$w = -\Delta PV.$$

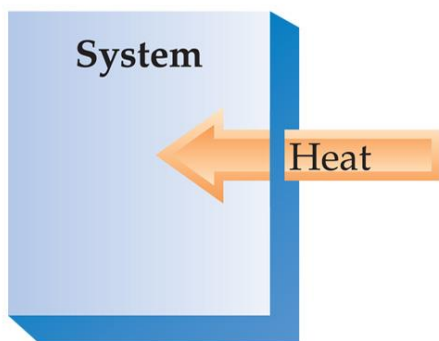
heat flow (q) = H (enthalpy) of system.

And: H is a state function, so q is also.

but only in the right conditions

Endothermic vs. Exothermic

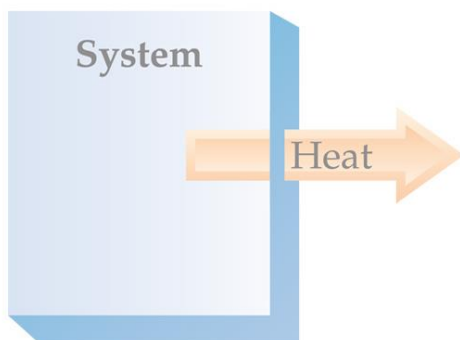
Surroundings



$\Delta H > 0$
(Endothermic)

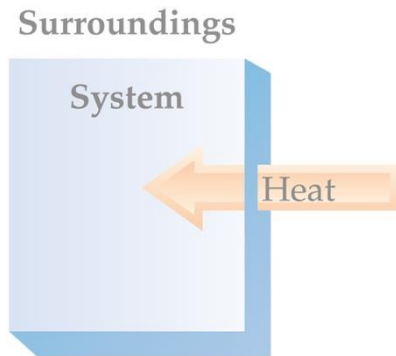
- A process is *endothermic* when ΔH is positive.

Surroundings

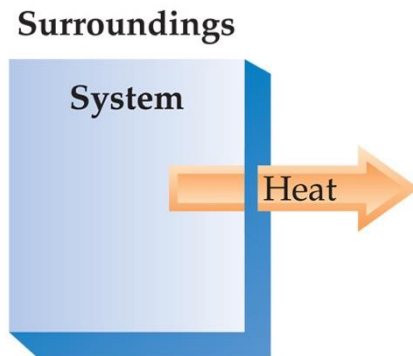


$\Delta H < 0$
(Exothermic)

Endothermicity and Exothermicity



$\Delta H > 0$
(Endothermic)



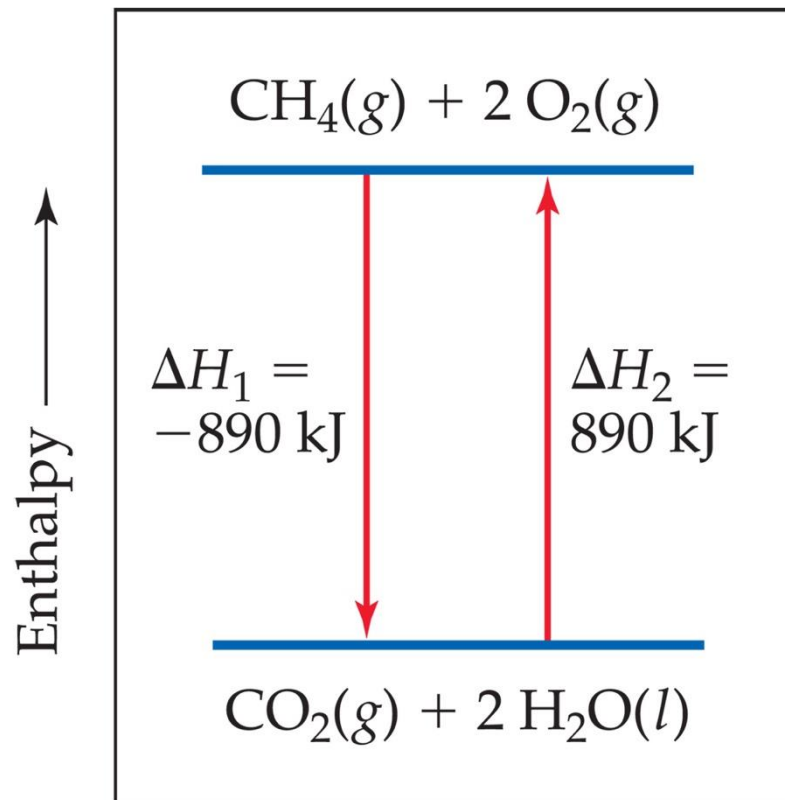
$\Delta H < 0$
(Exothermic)

- A process is *endothermic* when ΔH is **positive**.
- A process is *exothermic* when ΔH is **negative**.

Enthalpies of Reaction

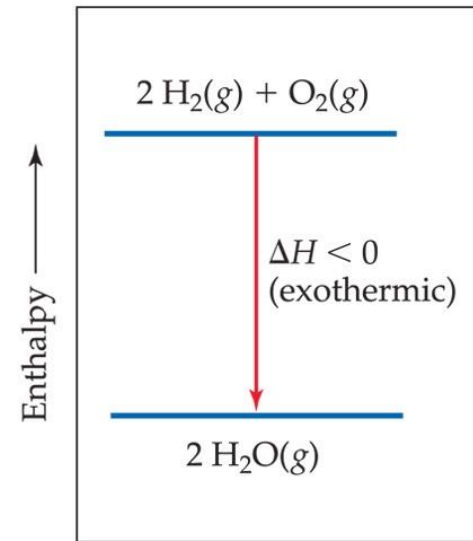
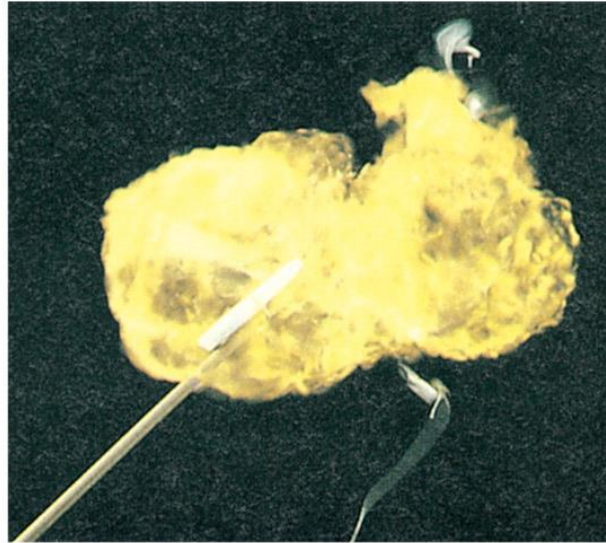
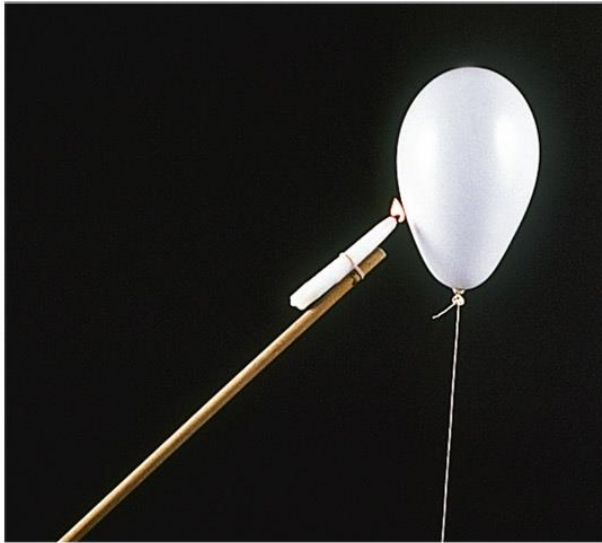
The *change* in enthalpy, ΔH , is the enthalpy of the products minus the enthalpy of the reactants:

$$\Delta H = H_{\text{products}} - H_{\text{reactants}}$$



Enthalpies of Reaction

This quantity, ΔH , is called the enthalpy of reaction, or the **heat** of reaction.



Reaction Enthalpy summary

1. Enthalpy is an *extensive* property.
2. ΔH for a reaction in the **forward** direction is **equal** in size, but **opposite** in **sign**, to ΔH for the **reverse** reaction.

$$\Delta H_{\text{(forward)}} = -\Delta H_{\text{(reverse)}}$$

1. ΔH for a reaction depends on the *state* of the products and reactants.

Enthalpy of reaction example

Consider the reaction:



a. What is the enthalpy change for decomposition of **2 moles** of KClO_3 ?

A. -45 kJ

B. -89.4 kJ

C. -180 kJ

D. 180 kJ

D. None of the above

Enthalpy of reaction example

Consider the reaction:



a. What is the enthalpy change for formation of **0.855 moles of O₂**?

Enthalpy of reaction example

Consider the reaction:



2 mol 2 mol 3 mol -89.4 kJ of heat released.

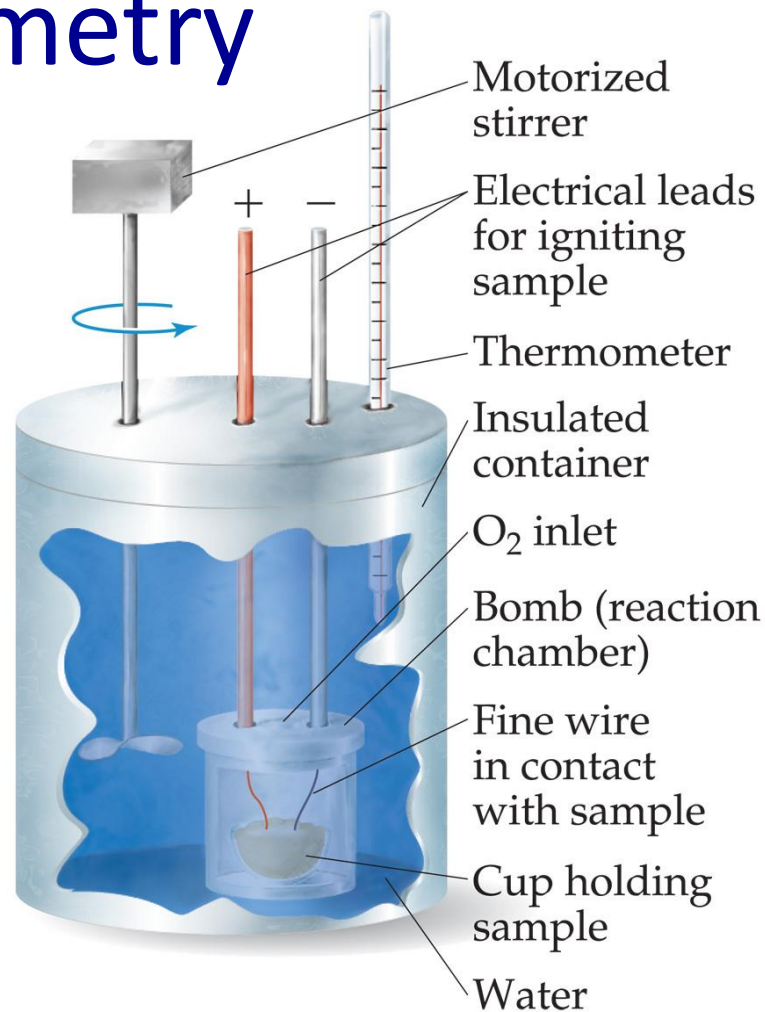
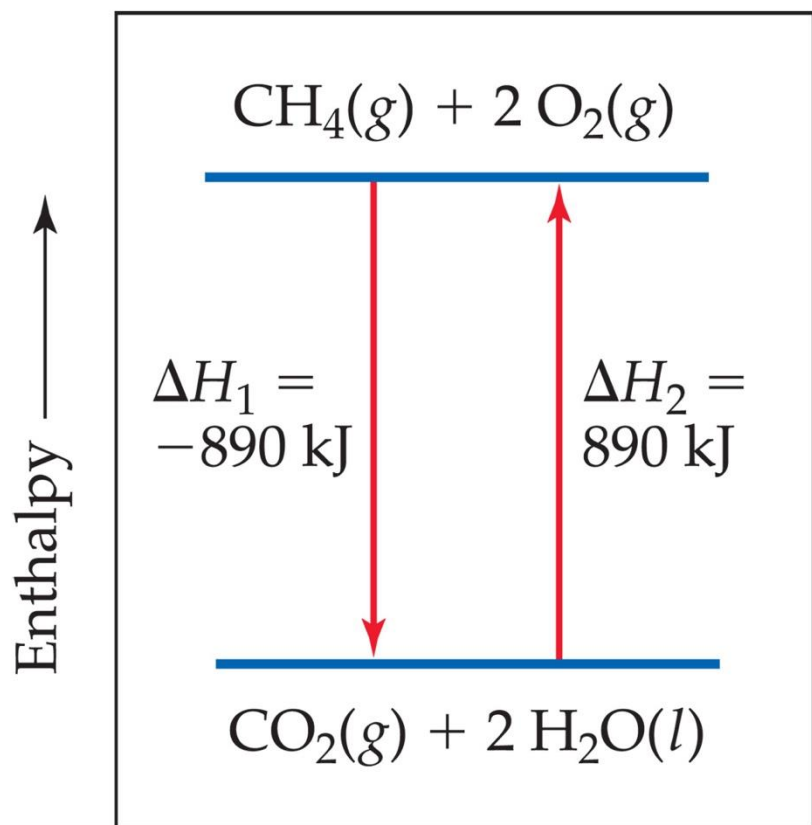
What is the enthalpy change for formation of 0.855 moles of O_2 ?



0.855 mol

$$\Delta H = -89.4 \text{ kJ}/(3 \text{ mol O}_2)(.855 \text{ mol O}_2) = -25.5 \text{ kJ}$$

Calorimetry



we measure ΔH through calorimetry, the measurement of heat flow.

Heat Capacity and Specific Heat

- **heat capacity:** amount of E required to raise the temperature of **an object, container, thing, etc.** by 1 K
- **specific heat:** amount of E required to raise the temperature of **1 g** of a substance by **1 K**.

Heat Capacity and Specific Heat

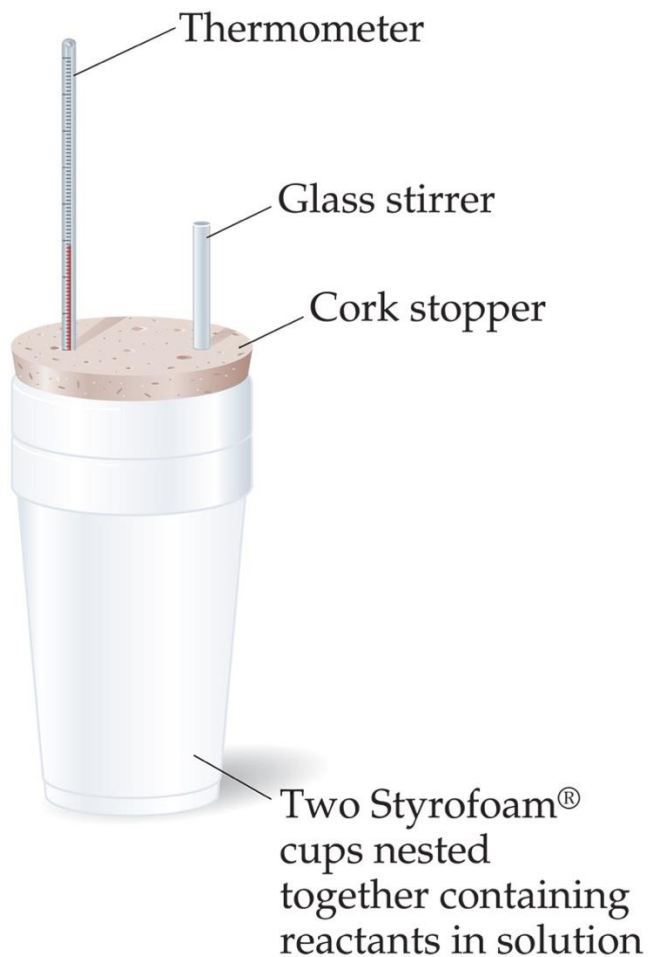
Specific heat is:

$$\text{Specific heat} = \frac{\text{heat transferred}}{\text{mass} \times \text{temperature change}}$$

$$s = \frac{q}{m \Delta T}$$

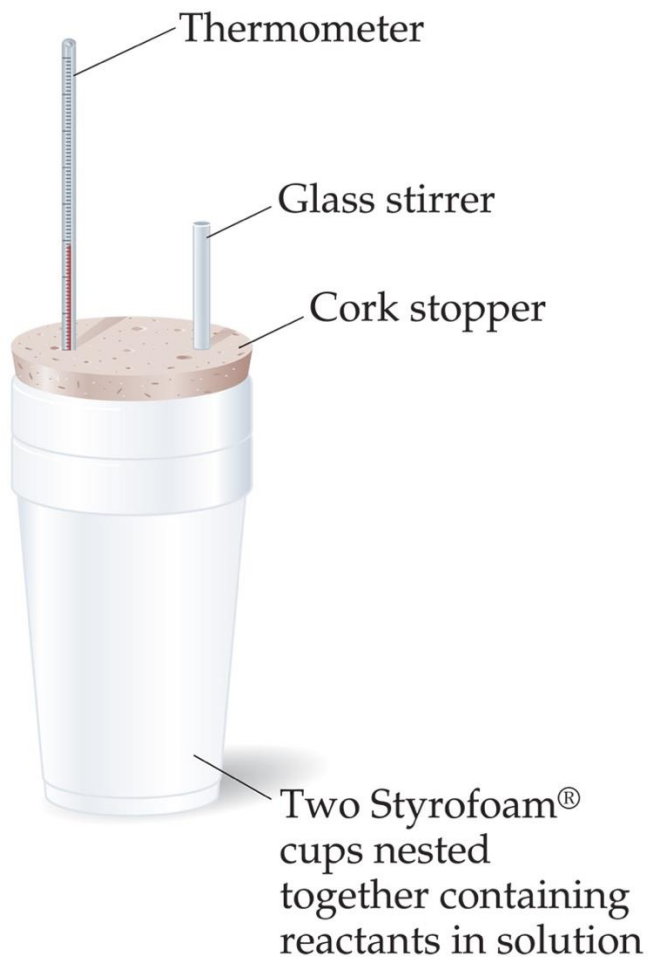
$$sm\Delta T = q$$

Constant Pressure Calorimetry



indirectly measure the heat change for the system by measuring the **heat change for the water** in the calorimeter.

Constant Pressure Calorimetry



s for water is 4.184 J/g-K
we can measure ΔH for
the reaction by measuring
the heat change of water:

$$q = m \times s \times \Delta T$$

$m = \text{mass}$

$s = \text{specific heat}$

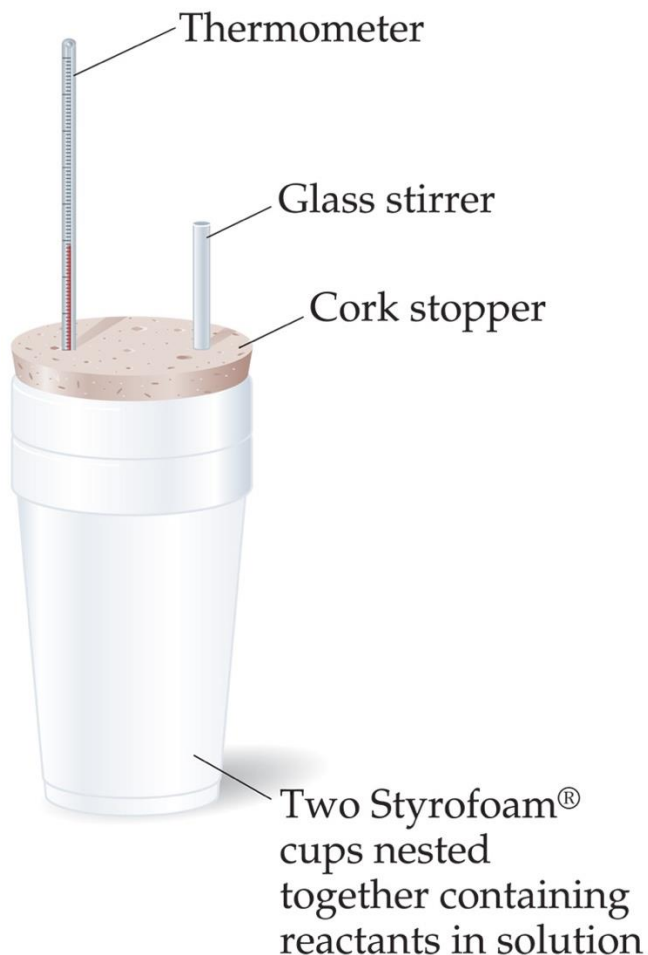
Constant Pressure Calorimetry

s for water is 4.184 J/g-K we can measure ΔH for the reaction by measuring the heat change of water:

$$q = m \times s \times \Delta T$$

$m = \text{mass}$

$s = \text{specific heat}$



How much heat was added to 100 g of water if the T rose from 20° C to 25° C ?

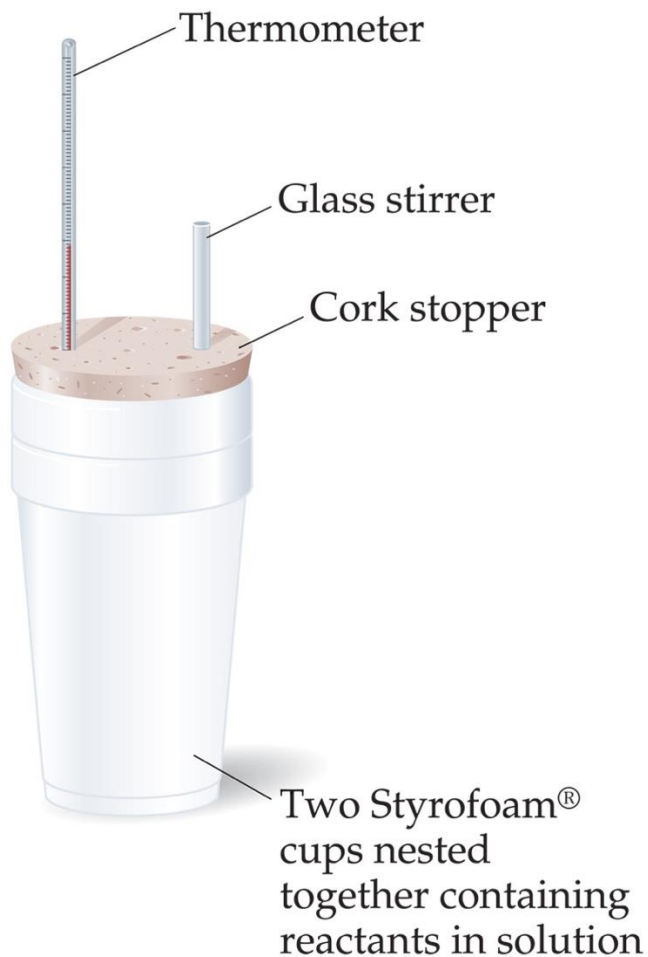
A. 20 J

B. 40 J

C. 400 J

D. 2000 J

Example



- When a **3.88 g** sample of solid ammonium nitrate dissolves in **60.0 g** of water in a coffee cup calorimeter, the temperature drops from **23.0 ° C** to **18.4 ° C**. (a) Calculate ΔH (in kJ/mol ammonium nitrate) for the solution process. Assume that the specific heat is constant and = $4.184 \text{ J/}^\circ \text{ C g}$. **$s \text{ of solution} = s_w$**
- (b) Is this process endothermic or exothermic?

Example

When a 3.88 g sample of solid ammonium nitrate dissolves in 60.0 g of water in a coffee cup calorimeter, the temperature drops from 23.0° C to 18.4° C. (a) Calculate ΔH (in kJ/mol ammonium nitrate) for the solution process. Assume that the specific heat is constant and = 4.184 J/g° C. (b) Is this process endothermic or exothermic?

Reaction:



gr 3.88 g

MW 80.04 g/mol

#Mol 3.88 g/80.04 g/mol = 0.0484 mol

Mass of solution = 3.88 g + 60 g = 63.88 g.

System: Solid AmNO₃

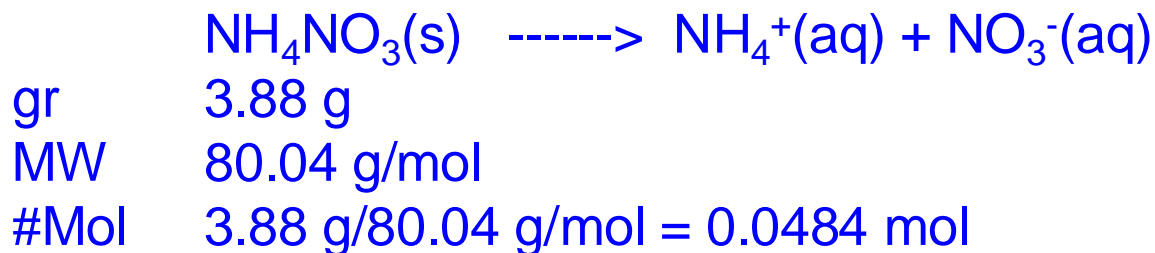
Surroundings: Solution (water AND dissolved NH₄NO₃)

Why? You are cooling BOTH the water AND the dissolved NH₄NO₃

Example

When a 3.88 g sample of solid ammonium nitrate dissolves in 60.0 g of water in a coffee cup calorimeter, the temperature drops from 23.0 ° C to 18.4 ° C. (a) Calculate ΔH (in kJ/mol ammonium nitrate) for the solution process. Assume that the specific heat is constant and = 4.184 J/g ° C. (b) Is this process endothermic or exothermic?

Reaction:



Mass of solution = 3.88 g + 60 g = 63.88 g.

System: Solid AmNO₃

Surroundings: Solution (water AND NH₄NO₃)

$q = sm\Delta T$. (s= specific heat, m= mass $\Delta T=T$ change).

$q = s(\text{J/g}^\circ \text{C})m(\text{grams})(T_{\text{final}} - T_{\text{initial}})$

$q_{\text{solution}} = 4.184(\text{J/g}^\circ \text{C})(63.88 \text{ g})(18.4^\circ \text{C} - 23.0^\circ \text{C}) = -1229 \text{ J}$

$q_{\text{water}} = -q_{\text{ammonium nitrate}} = +1229 \text{ J}$

b. Is the process endothermic or exothermic?

A. Endothermic

B Exothermic

Example

When a 3.88 g sample of solid ammonium nitrate dissolves in 60.0 g of water in a coffee cup calorimeter, the temperature drops from 23.0 ° C to 18.4 ° C. (a) Calculate ΔH (in kJ/mol ammonium nitrate) for the solution process. Assume that the specific heat is constant and = 4.184 J/g ° C. (b) Is this process endothermic or exothermic?

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$$q = s(\text{specific heat})m(\text{mass})\Delta T$$

$$q = s(\text{J/g}^\circ \text{C})m(\text{grams})(T_{\text{final}} - T_{\text{initial}})$$

$$q_{\text{solution}} = 4.184(\text{J/g}^\circ \text{C})(63.88 \text{ g})(18.4^\circ \text{C} - 23.0^\circ \text{C}) = -1229 \text{ J}$$

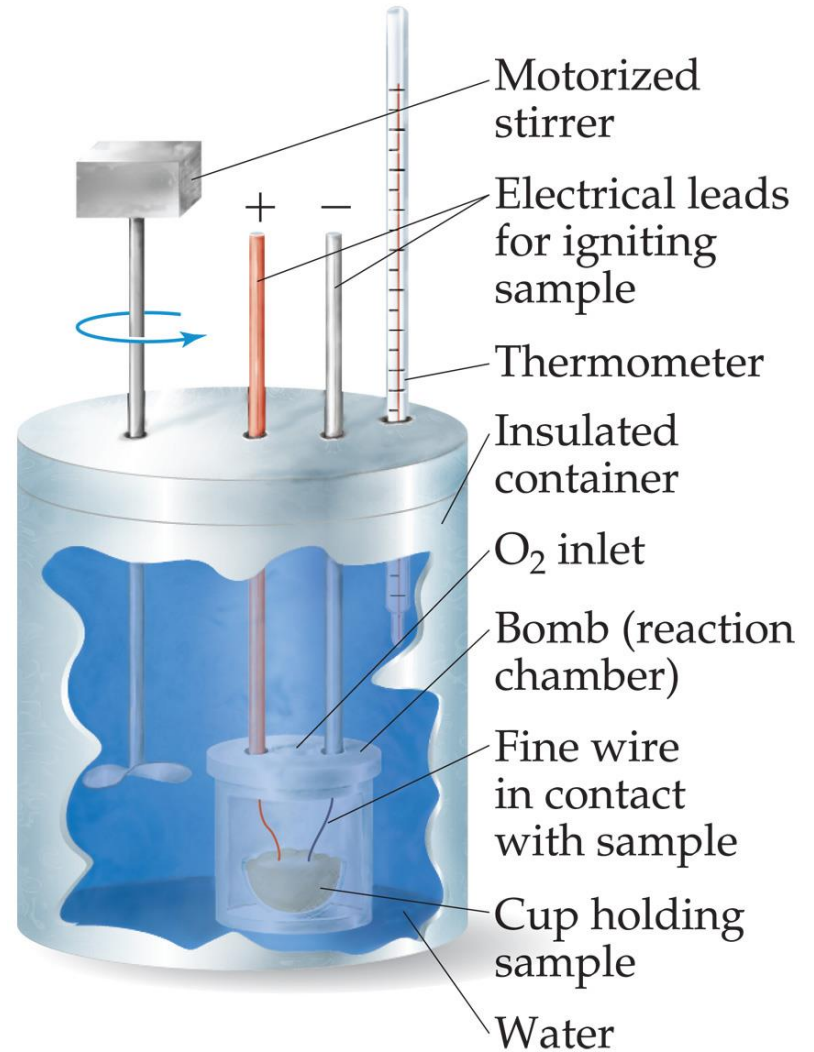
$$q_{\text{water}} = -q_{\text{ammonium nitrate}} = +1229 \text{ J}$$

$$\Delta H(\text{per mol NH}_4\text{NO}_3) = 1.229 \text{ kJ}/.0484 \text{ mol} = 25.39 \text{ kJ/mol}$$

(b) **Endothermic**

Bomb Calorimetry

Reactions can be carried out separated from the water in a “bomb,” such as this one,
And still measure the heat absorbed by the water.



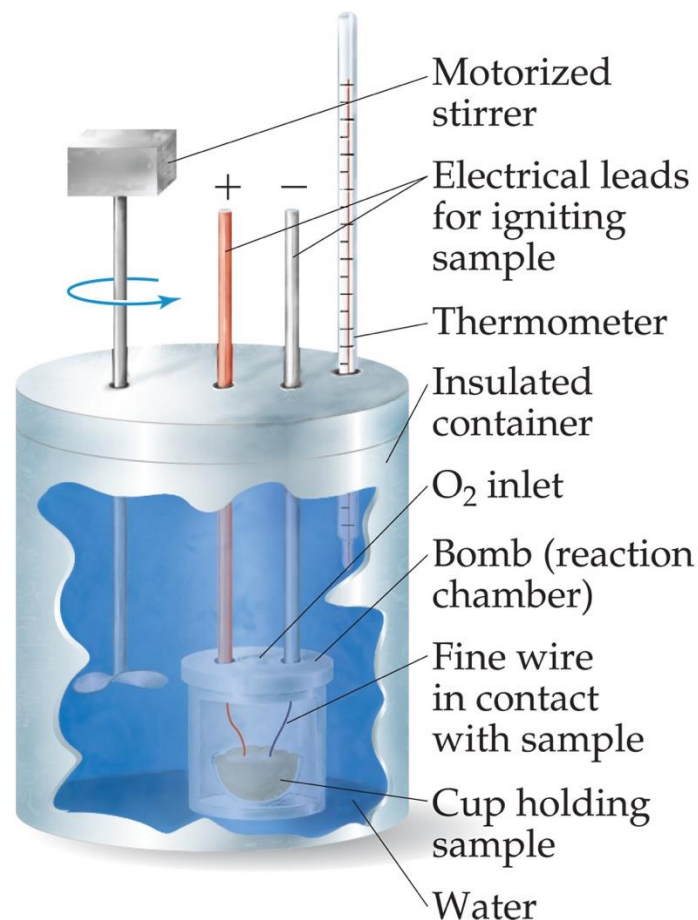
Bomb Calorimetry

- Because the volume in the bomb calorimeter is constant, what is measured is really the ΔE , not ΔH .

- For most reactions,

$\Delta E \approx \Delta H$

- *Why?*



Bomb Calorimetry

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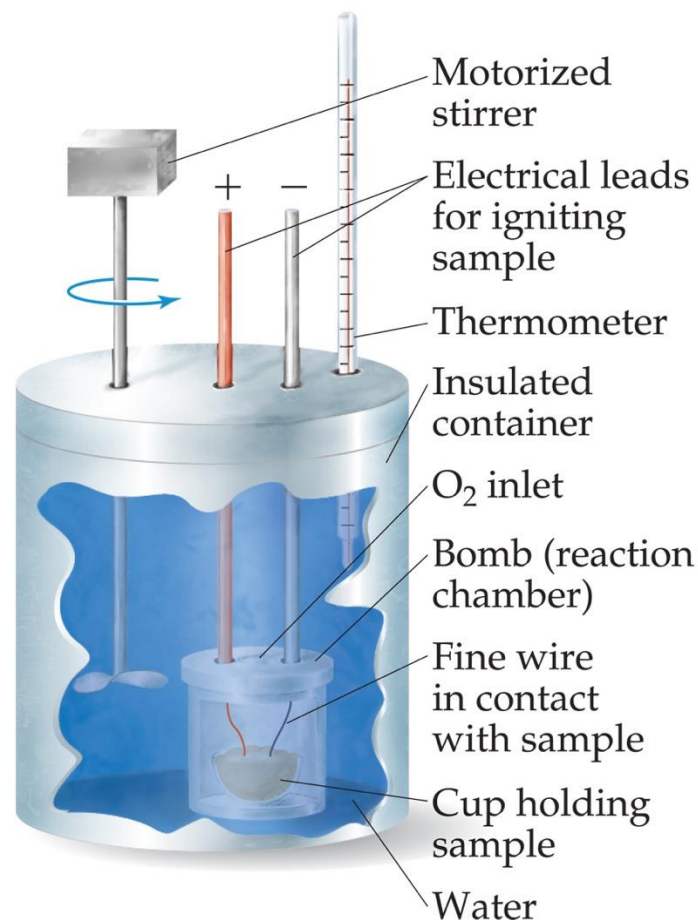
$\Delta E \approx \Delta H$

- *Why?*

- *No work, no Volume change.*

$\Delta E = q + w \approx \Delta H = \Delta E + \Delta PV$

Both 0, no V or P change.



Bomb Calorimetry

$$H = E + PV$$

$$\Delta H = \Delta E + \Delta PV$$

In a bomb calorimeter, $\Delta V = 0$

For a process that doesn't evolve gas:

$\Delta P \approx 0$ as well.

$$\Delta H = \Delta E + \Delta PV = \Delta E$$

Example

- A 50 g sample of gasoline was burned by combustion (with excess oxygen) in a calorimeter with a heat capacity of 10 kJ/ °C. The temperature increased by 100 °C. Calculate the change in E per g of gasoline.

- $q_{\text{surroundings}} = C\Delta T = 10 \text{ kJ/}^{\circ}\text{C}(100^{\circ}\text{C}) = 1000 \text{ kJ}$

- $q_{\text{surroundings}} = -q_{\text{system}}$

- $q_{\text{system}} = -1000$

- $-1000 \text{ kJ}/50\text{g} = -20 \text{ kJ/g}$

- **Does $\Delta E = \Delta H$ in this case?**

- **A. yes**

- **B. no**

Example

- A 50 g sample of gasoline was burned by combustion (with excess oxygen) in a calorimeter with a heat capacity of $10 \text{ kJ/}^\circ \text{C}$. The temperature increased 100°C . Calculate the change in E per g of gasoline.

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- $-1000 \text{ kJ}/50\text{g} = -20 \text{ kJ/g}$

- **Does $\Delta E = \Delta H$ in this case?**

- **NO! Combustion leads to evolution of gases (CO_2 gas Pressure **can't** stay constant in this case.**

Hess' s Law

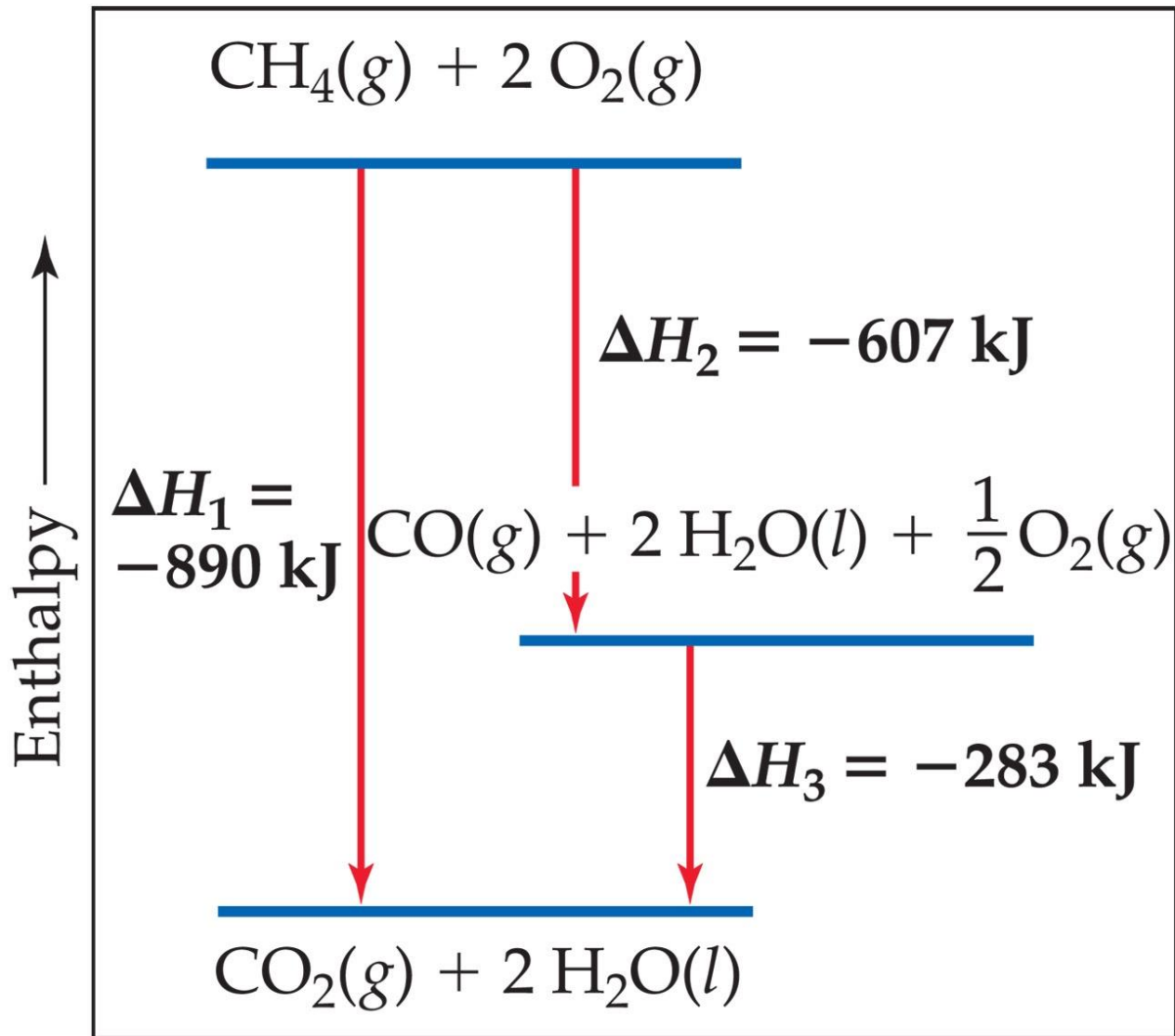
□ $\Delta H = \Delta E + \Delta PV = q_p$

□ ΔH is known for many reactions.

• measuring ΔH can be a pain

• *Can we estimate ΔH using ΔH values for other reactions?*

Hess' s Law



Yes!

Hess' s law:
states that:

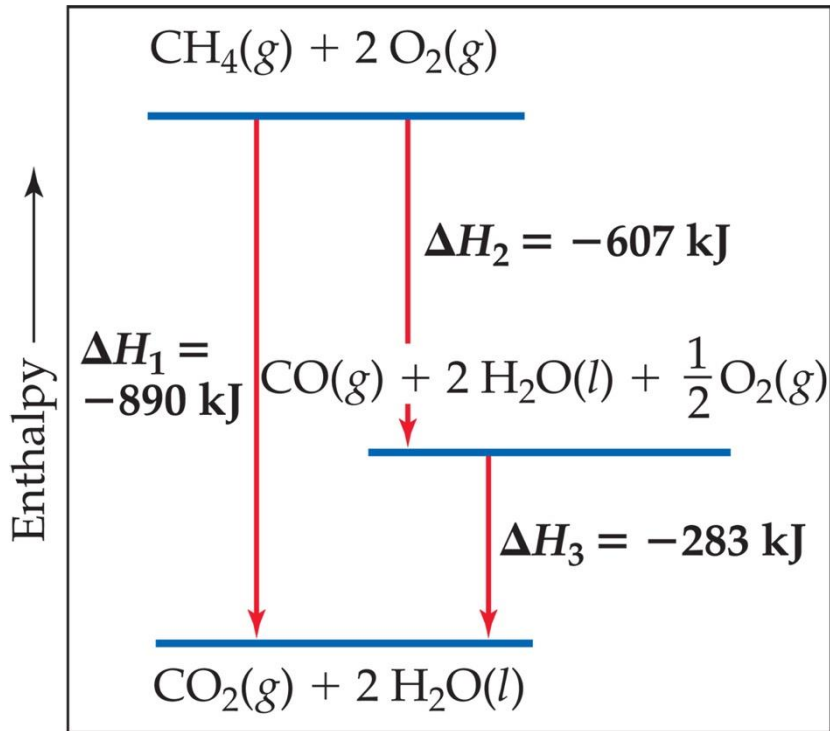
ΔH for the overall
reaction will be
equal to the sum
of the enthalpy
changes for the
individual steps.

Hess' s Law

Why?

Because ΔH is a state function,
and is pathway independent.

Only depends on initial state of
the reactants and the final
state of the products.



A series of chemical reactions is a
path.

start with the **same reactants**
end up with the **same products**,
 ΔH is the same.

because it's a state function.

Hess' s law, example:

- Given:
- $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \text{ ----> } 2\text{NO}(\text{g}) \quad \Delta\text{H} = 180.7 \text{ kJ}$
- $2\text{NO}(\text{g}) + \text{O}_2(\text{g}) \text{ ----> } 2\text{NO}_2(\text{g}) \quad \Delta\text{H} = -113.1 \text{ kJ}$
- $2\text{N}_2\text{O}(\text{g}) \text{ ----> } 2\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \quad \Delta\text{H} = -163.2 \text{ kJ}$
- use Hess' s law to calculate ΔH for the reaction:
- $\text{N}_2\text{O}(\text{g}) + \text{NO}_2(\text{g}) \text{ ----> } 3\text{NO}(\text{g})$

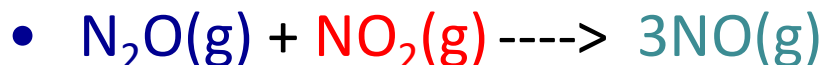
- How do you think through this?
- What do I need on **left** side of the arrow? $\text{N}_2\text{O}(\text{g})$ and $\text{NO}_2(\text{g})$
- What do I need on the **right** side of the arrow? $3\text{NO}(\text{g})$

Hess' s law, example:

• Given:



• use Hess' s law to calculate ΔH for the reaction:



Enthalpies of Formation ΔH_f

An enthalpy (heat) of formation, ΔH_f , is defined as the ΔH for the reaction in which a compound is made from its constituent elements in their **most stable** elemental forms.



- What is the heat of reaction given:



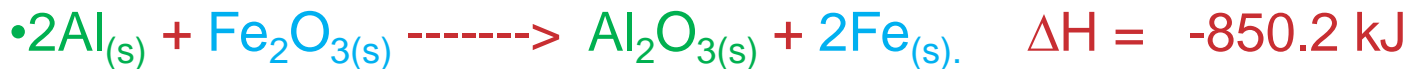
Enthalpies of Formation

An enthalpy (heat) of formation, ΔH_f , is defined as the ΔH for the reaction in which a compound is made from its constituent elements in their **most stable** elemental forms.

Example:



•What is the heat of reaction given:



Enthalpies of Formation

ΔH_f , is defined as the ΔH for the reaction in which a compound is made from its constituent elements in their ***most stable*** elemental forms.

In other words, a reaction like these:



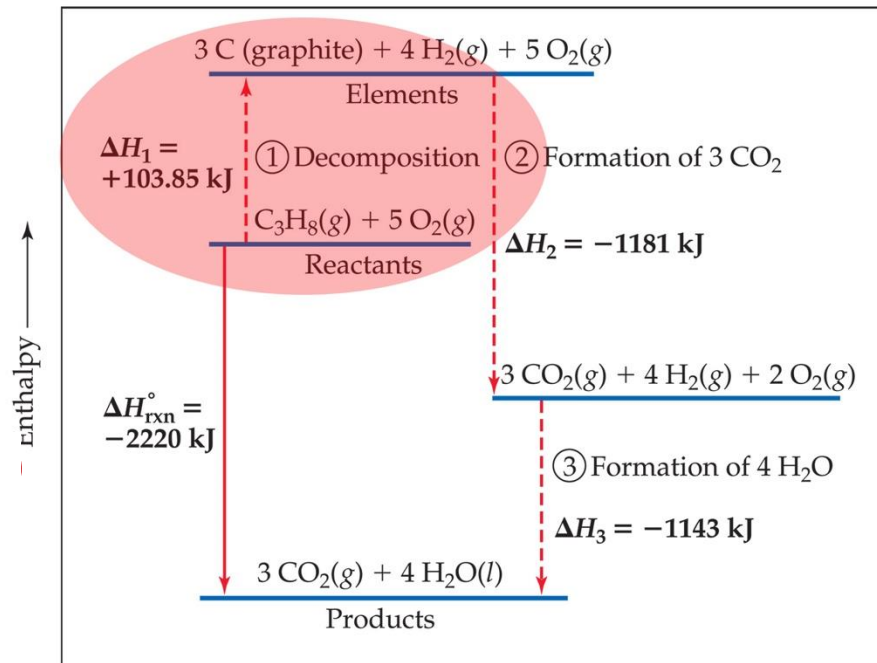
Elements in their most stable forms:

- $\text{H}_{2(g)}$, $\text{N}_{2(g)}$, $\text{O}_{2(g)}$, $\text{F}_{2(g)}$, $\text{Cl}_{2(g)}$
- $\text{Br}_{2(l)}$, $\text{I}_{2(s)}$
- All metals (except Hg): $\text{M}_{(s)}$
- $\text{Hg}_{(l)}$
- All other nonmetals: $\text{X}_{(s)}$ example: $\text{S}_{(s)}$, $\text{Si}_{(s)}$

Calculation of ΔH



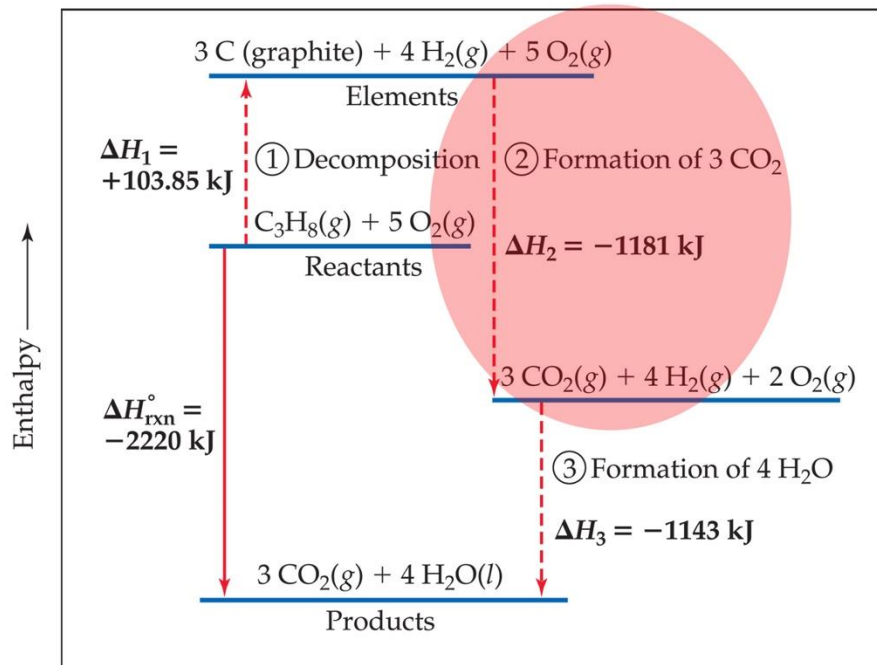
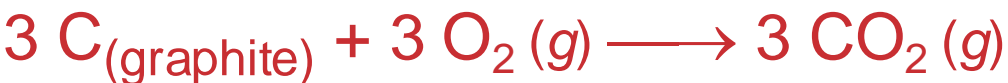
- Imagine this as occurring in 3 steps:



Calculation of ΔH



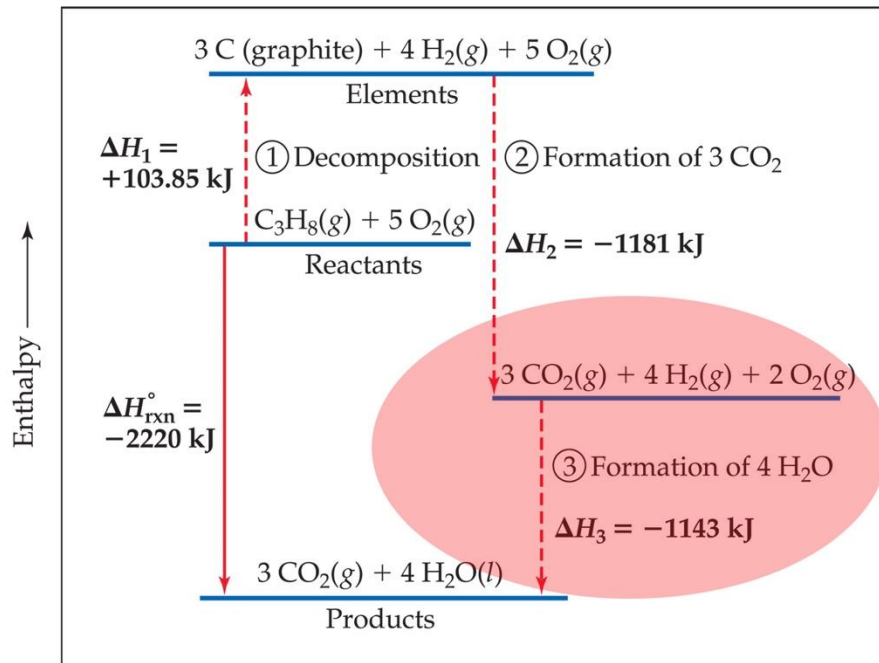
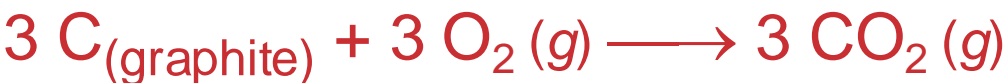
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Calculation of ΔH



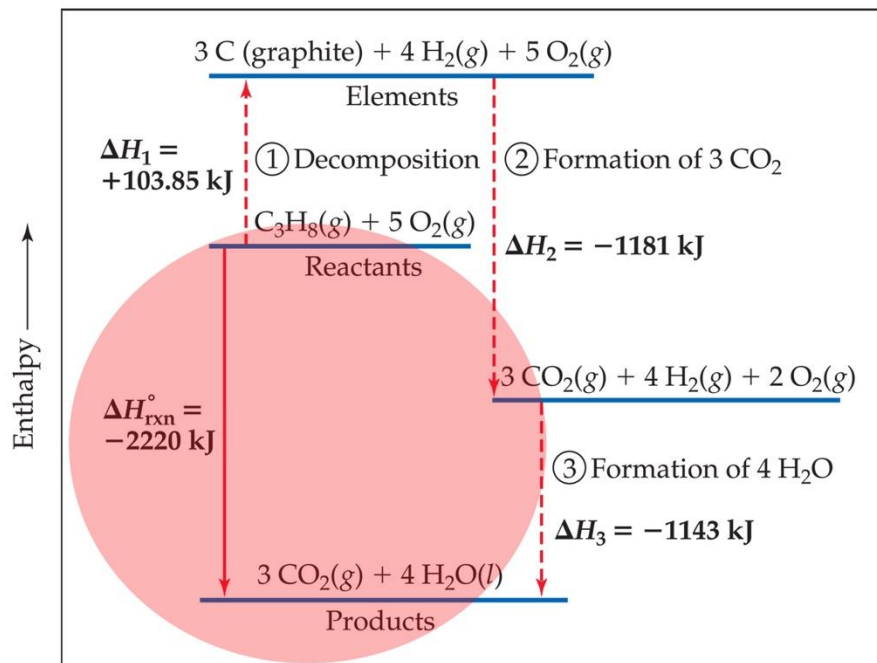
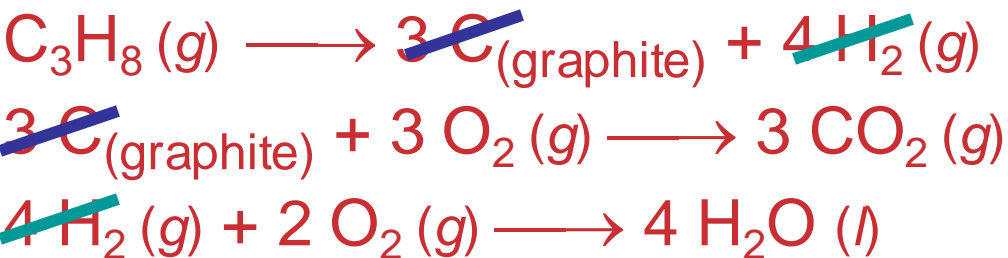
- Imagine this as occurring in 3 steps:



Calculation of ΔH



- The sum of these equations is:



Make each reactant or product from its elements
 This is called the heat of formation of a compound

Calculation of ΔH

Another way of saying Hess' s law

$$\Delta H = \sum n \Delta H_{f(\text{products})}^{\circ} - \sum m \Delta H_{f(\text{reactants})}^{\circ}$$

where n and m are the stoichiometric coefficients.

Why is this useful?

Because there are humungus tables of ΔH_f

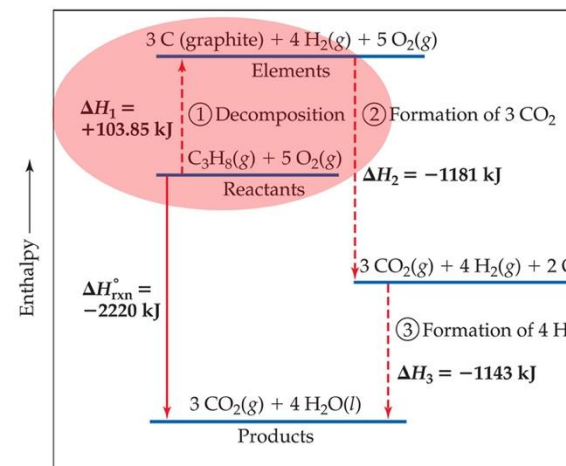
Standard Enthalpies of Formation

Standard enthalpies of formation, ΔH_f° , are measured under standard conditions (25° C and 1.00 atm pressure).

Substance	Formula	ΔH_f° (kJ/mol)	Substance	Formula	ΔH_f° (kJ/mol)
Acetylene	$C_2H_2(g)$	226.7	Hydrogen chloride	$HCl(g)$	-92.30
Ammonia	$NH_3(g)$	-46.19	Hydrogen fluoride	$HF(g)$	-268.60
Benzene	$C_6H_6(l)$	49.0	Hydrogen iodide	$HI(g)$	25.9
Calcium carbonate	$CaCO_3(s)$	-1207.1	Methane	$CH_4(g)$	-74.80
Calcium oxide	$CaO(s)$	-635.5	Methanol	$CH_3OH(l)$	-238.6
Carbon dioxide	$CO_2(g)$	-393.5	Propane	$C_3H_8(g)$	-103.85
Carbon monoxide	$CO(g)$	-110.5	Silver chloride	$AgCl(s)$	-127.0
Diamond	$C(s)$	1.88	Sodium bicarbonate	$NaHCO_3(s)$	-947.7
Ethane	$C_2H_6(g)$	-84.68	Sodium carbonate	$Na_2CO_3(s)$ 1	-130.9
Ethanol	$C_2H_5OH(l)$	-277.7	Sodium chloride	$NaCl(s)$	-410.9
Ethylene	$C_2H_4(g)$	52.30	Sucrose	$C_{12}H_{22}O_{11}(s)$	-2221
Glucose	$C_6H_{12}O_6(s)$	-1273	Water	$H_2O(l)$	-285.8
Hydrogen bromide	$HBr(g)$	-36.23	Water vapor	$H_2O(g)$	-241.8

Calculation of ΔH

- Calculate ΔH using the table:
- $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$



Substance	Formula	ΔH_f° (kJ/mol)	Substance	Formula	ΔH_f° (kJ/mol)
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What's the heat of formation of $O_2(g)$?

- A. Can't tell it's not in the table
- B. -200 kJ/mol
- C. 0

Calculation of ΔH

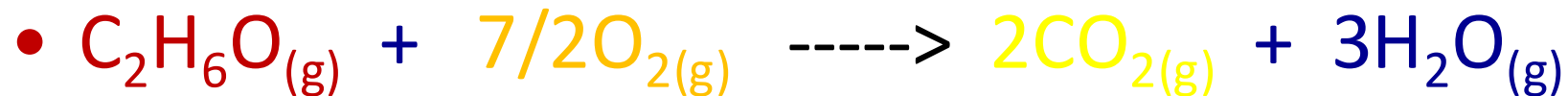


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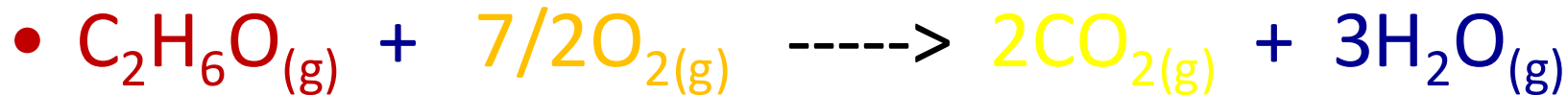
$$\begin{aligned} \Delta H &= [3(\Delta H_f \text{CO}_{2(g)}) + 4(\Delta H_f \text{H}_2\text{O}_{(l)})] - [(\Delta H_f \text{C}_3\text{H}_{8(g)}) + (5\Delta H_f \text{O}_{2(g)})] \\ &= [3(-393.5 \text{ kJ}) + 4(-285.8 \text{ kJ})] - [(-103.85 \text{ kJ}) + 5(0)] \\ &= [-1180.5 \text{ kJ} + (-1143.2 \text{ kJ})] - [(-103.85 \text{ kJ}) + 0 \text{ kJ}] \\ &= [-2323.7 \text{ kJ}] - [-103.85 \text{ kJ}] \\ &= -2219.9 \text{ kJ} \end{aligned}$$

Another example

a. Write a balanced equation for the combustion of ethanol to produce water vapor and carbon dioxide.



Another example



b. Calculate the standard enthalpy change for the reaction, assuming $\text{H}_2\text{O}(g)$ as a product.

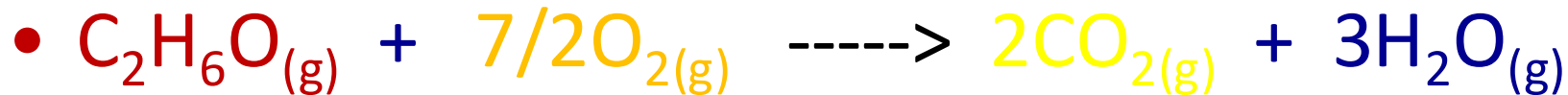
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$$\Delta H = [2(\Delta H_f \text{CO}_{2(g)}) + 3(\Delta H_f \text{H}_2\text{O}_{(g)})] - [(\Delta H_f \text{C}_2\text{H}_6\text{O}_{(g)}) + (3\Delta H_f \text{O}_{2(g)})]$$

$$= [2(-393.5 \text{ kJ}) + 3(-285.8 \text{ kJ})] - [(-277.7 \text{ kJ}) + 5(0)] = -1234.7 \text{ kJ}$$

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a. Write a balanced equation for the combustion of ethanol to produce water vapor.



b. Calculate the standard enthalpy change for the reaction, assuming $\text{H}_2\text{O}(g)$ as a product.

$$\begin{aligned}\Delta H &= [2(\Delta H_f \text{CO}_{2(g)}) + 3(\Delta H_f \text{H}_2\text{O}_{(g)})] - [(\Delta H_f \text{C}_2\text{H}_6\text{O}_{(g)}) + (3\Delta H_f \text{O}_{2(g)})] \\ &= [2(-393.5 \text{ kJ}) + 3(-285.8 \text{ kJ})] - [(-277.7 \text{ kJ}) + 5(0)] = -1234.7 \text{ kJ}\end{aligned}$$

c. Calculate the heat produced per liter of ethanol by combustion of ethanol under constant pressure. Ethanol has a density of 0.789 g/mL.

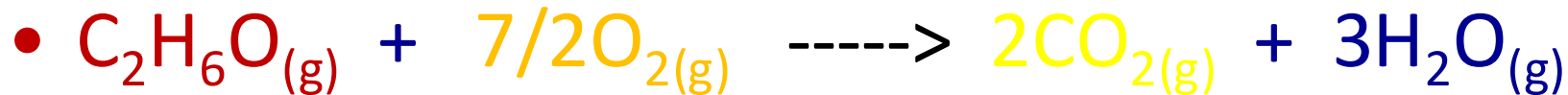
$$1000 \text{ mL}(0.789\text{g/mL}) = 789 \text{ g ethanol in 1 liter.}$$

$$\text{M.M. ethanol: } 2\text{C} + 6\text{H} + 1\text{O} = 24 + 6 + 16 = 46 \text{ g/mol}$$

$$789 \text{ g/L ethanol})(1\text{mole}/46 \text{ g ethanol}) = 17.15 \text{ moles/L ethanol.}$$

$$\text{Heat from 1 liter of ethanol} = (17.15 \text{ moles/L})(-1234.7 \text{ kJ/mole}) = -21,178 \text{ kJ/L}$$

Another example



b. Calculate the standard enthalpy change for the reaction, assuming H₂O(g) as a product.

$$\begin{aligned}\Delta H &= [2(\Delta H_f \text{CO}_{2(g)}) + 3(\Delta H_f \text{H}_2\text{O}_{(g)})] - [(\Delta H_f \text{C}_2\text{H}_6\text{O}_{(g)}) + (3\Delta H_f \text{O}_{2(g)})] \\ &= [2(-393.5 \text{ kJ}) + 3(-285.8 \text{ kJ})] - [(-277.7 \text{ kJ}) + 5(0)] = -1234.8 \text{ kJ}\end{aligned}$$

c. Calculate the heat produced per liter of ethanol by combustion of ethanol under constant pressure. Ethanol has a density of 0.789 g/mL.

$$\text{Heat from 1 liter of ethanol} = (17.15 \text{ moles/L})(-1234.7 \text{ kJ/mole}) = -21178 \text{ kJ/L}$$

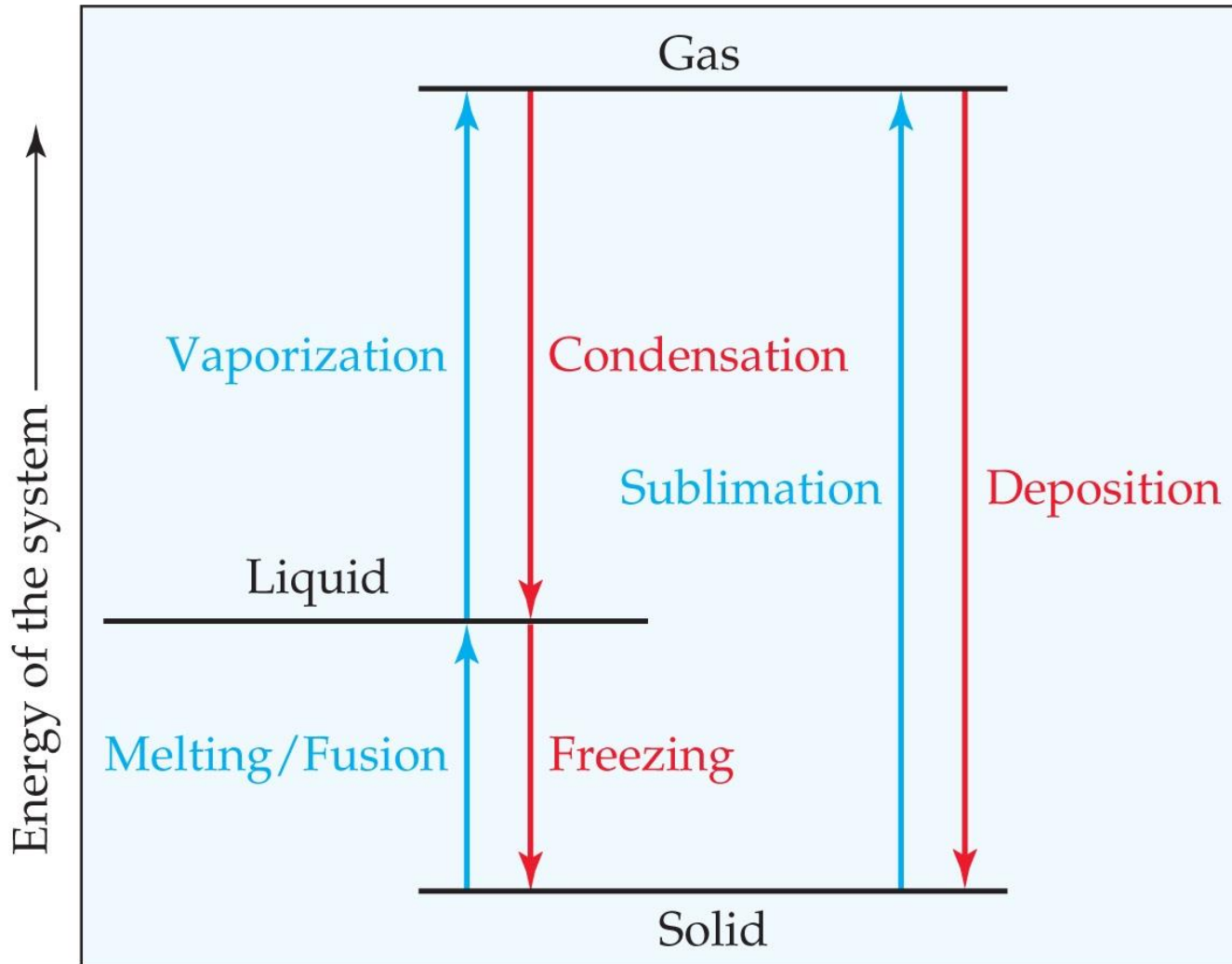
d. Calculate the mass of CO₂ produced per kJ of heat emitted. Express your answer using five significant figures.

$$1234.8 \text{ kJ}/(2 \text{ moles CO}_2)$$

$$\text{M.M. CO}_2 = 12 + (16)2 = 44 \text{ g/mol}$$

$$88.019 \text{ g CO}_2/ 1234.78 \text{ kJ} = .071284 \text{ g/kJ CO}_2$$

Phase Changes



Conversion from one state of matter to another is a **phase change**.

Energy is either added or released in a phase change.

— Endothermic process (energy added to substance)

— Exothermic process (energy released from substance)

Why is E (Q,H) transferred in phase change?

• Atoms/molecules

- **stick** to each other in **liquid/solid** always stick more in solid.
- Atoms/molecules **don't touch** in **gas**.

• Solid to liquid

- Heat added to unstick.

• Liquid to gas,

- E is needed to pull the atoms/molecules **completely** away from each other.

Strength of intermolecular attractions increasing



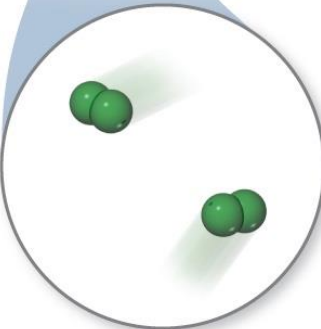
Gas



Liquid

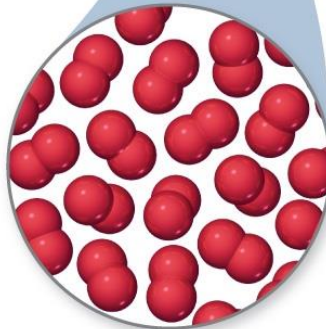


Crystalline solid



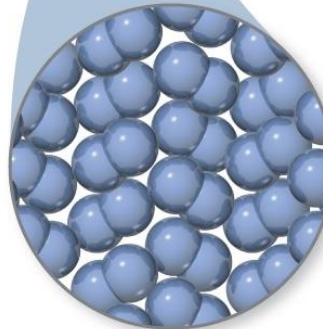
Chlorine, Cl₂

Particles far apart; possess complete freedom of motion



Bromine, Br₂

Particles are closely packed but randomly oriented; retain freedom of motion; rapidly change neighbors



Iodine, I₂

Particles are closely packed in an ordered array; positions are essentially fixed

Why is E (Q,H) transferred in phase change?

Strength of intermolecular attractions increasing

• Solid to liquid

• Liquid to gas,

• Which phase transfer will **absorb** the most energy per gram?

A. solid to liquid

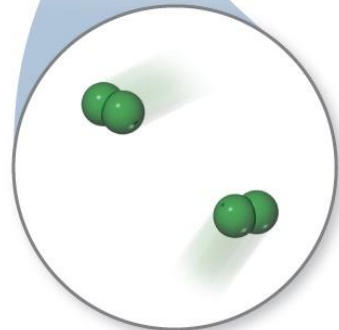
B. Liquid to gas

C. liquid to solid

D. gas to liquid



Gas

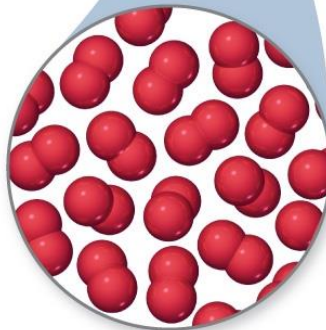


Chlorine, Cl_2

Particles far apart; possess complete freedom of motion



Liquid

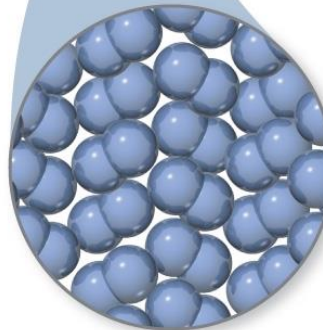


Bromine, Br_2

Particles are closely packed but randomly oriented; retain freedom of motion; rapidly change neighbors



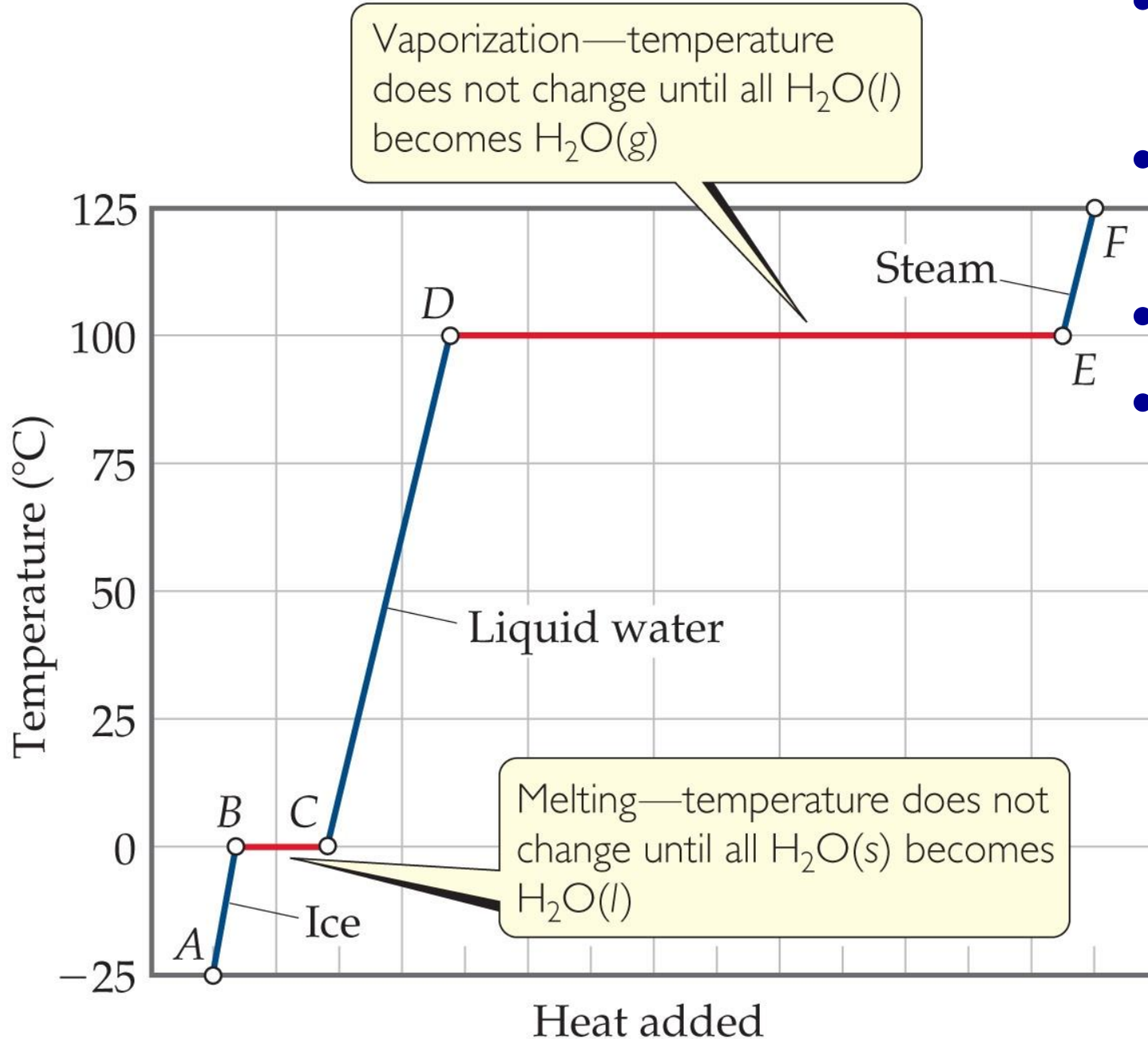
Crystalline solid



Iodine, I_2

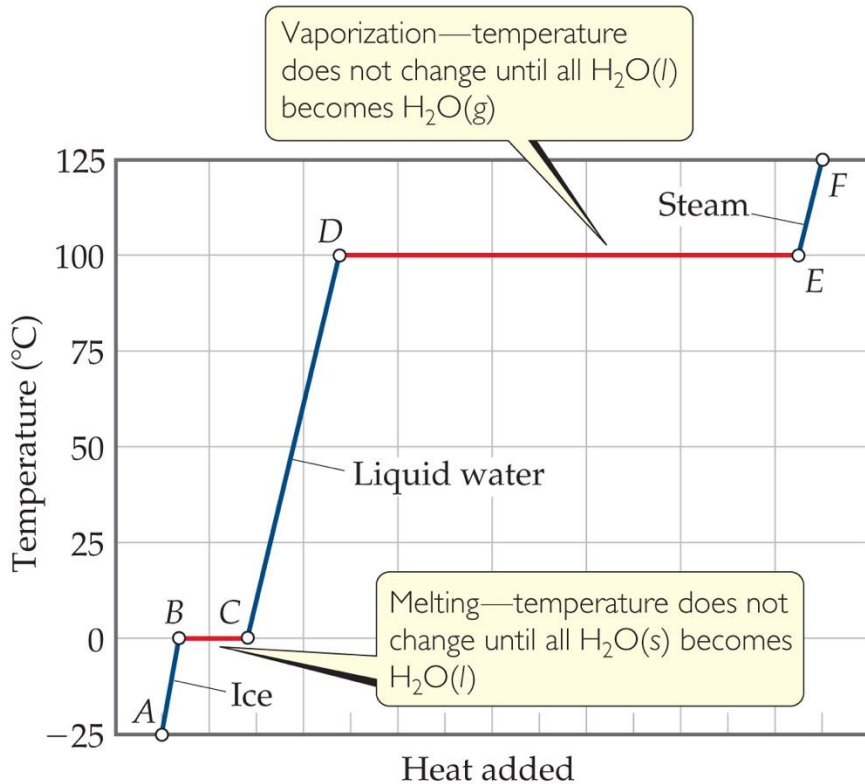
Particles are closely packed in an ordered array; positions are essentially fixed

Heating Curves



- A plot of T vs. q
- Within a phase:
 - $q = ms\Delta T$
 - The **temperature** of the substance **does not rise** during a phase change.

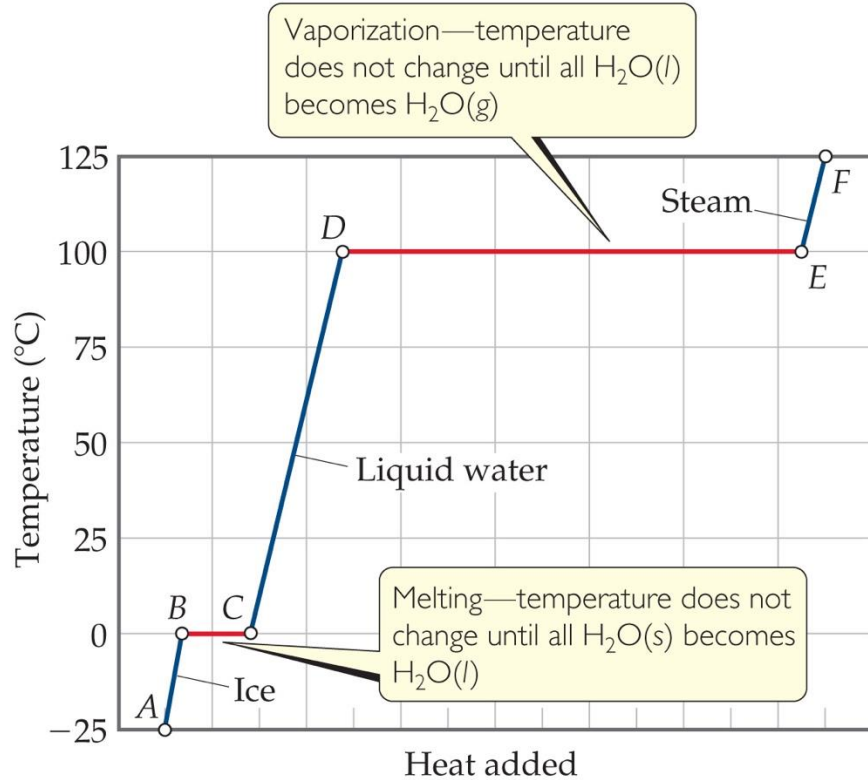
Heating Curves



- A plot of T vs. q
- Within a phase:
 - $q = ms\Delta T$
- During melting:
 - $q = \Delta H_{\text{FUS}}m$
- During boiling:
 - $q = \Delta H_{\text{vap}}m$

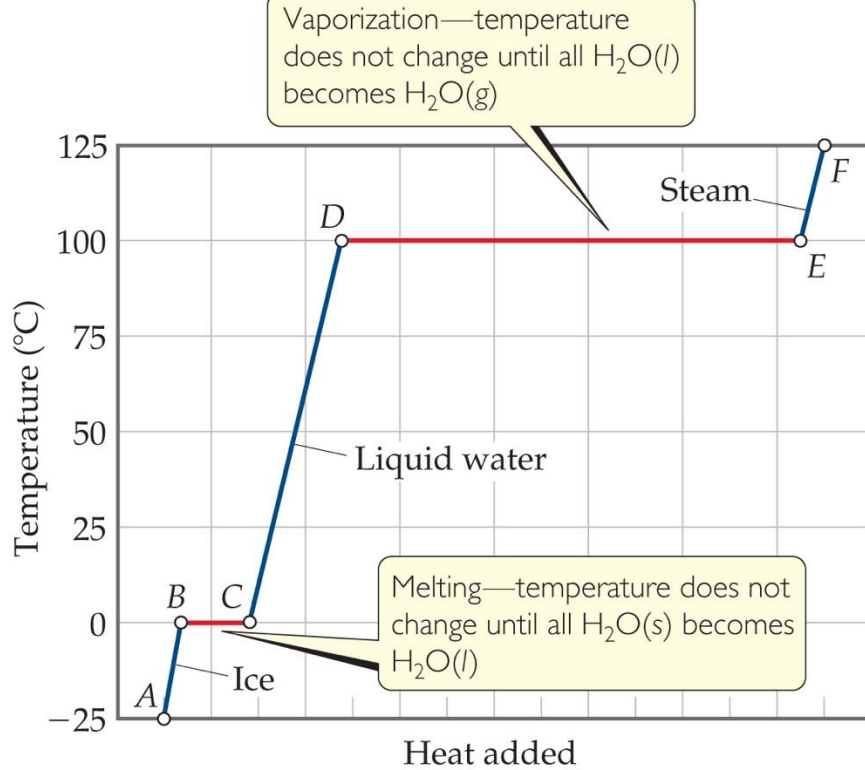
- For the phase changes, the product of mass and the heat of fusion or vaporization is heat.

Example:



- Calculate q for taking 10g ice from A to F:

Example:



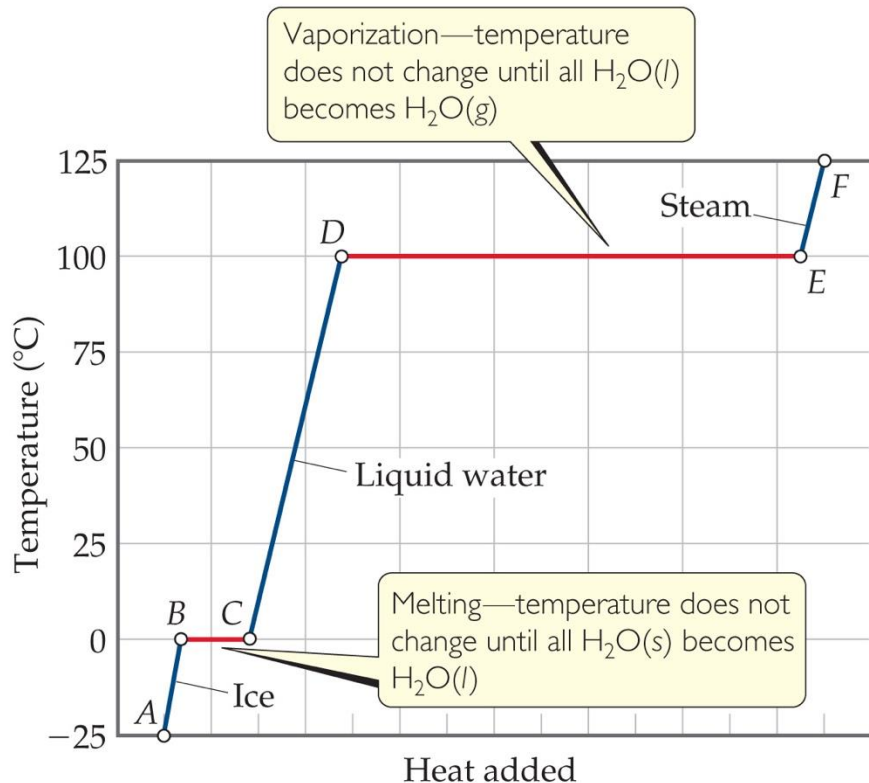
$$\Delta H_{\text{Fus}} \text{ water: } 334 \text{ J/g}$$

$$\Delta H_{\text{vap}} \text{ water: } 2444 \text{ J/g}$$

$$S_{\text{liquid water}} \quad 4.184 \text{ J/g}^\circ \text{K}$$

$$S_{\text{ice}} \quad 2.05 \text{ J/gK}$$

$$S_{\text{water vapor}} \quad 2.00 \text{ J/gK}$$



Example:

$$\Delta H_{\text{Fus}} \text{ water: } 334 \text{ J/g}$$

$$\Delta H_{\text{vap}} \text{ water: } 2444 \text{ J/g}$$

$$S_{\text{liquid water:}} 4.184 \text{ J/g}^\circ \text{ K}$$

$$S_{\text{ice:}} 2.05 \text{ J/gK}$$

$$S_{\text{water vapor:}} 2.00 \text{ J/gK}$$

You have to calculate each step separately:

Calculate q for taking 10g ice from A to F:

$$\begin{aligned} \text{A} \rightarrow \text{B}: q &= m S_{\text{ice}} \Delta T \\ (10\text{g}) 2.05 \text{ J/gK} (25 \text{ K}) &= 512.5 \text{ J} \end{aligned}$$

$$\begin{aligned} \text{B} \rightarrow \text{C}: q &= (m) \Delta H_{\text{Fus}} \\ 10\text{g} (334 \text{ J/g}) &= 3340 \text{ J} \end{aligned}$$

$$\begin{aligned} \text{C} \rightarrow \text{D}: q &= m (S_{\text{liquid water}}) \Delta T \\ 10\text{g} (4.184 \text{ J/gK}) (100 \text{ K}) &= 4184 \text{ J} \end{aligned}$$

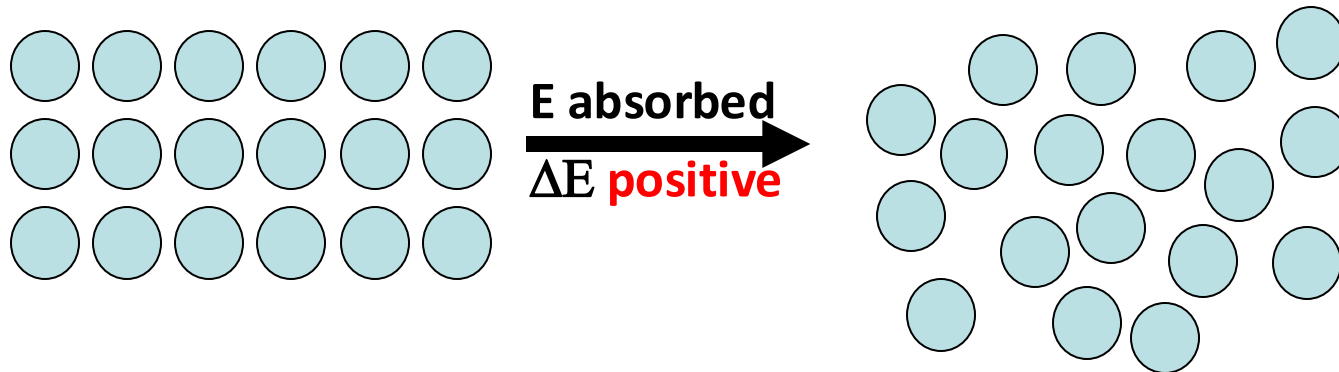
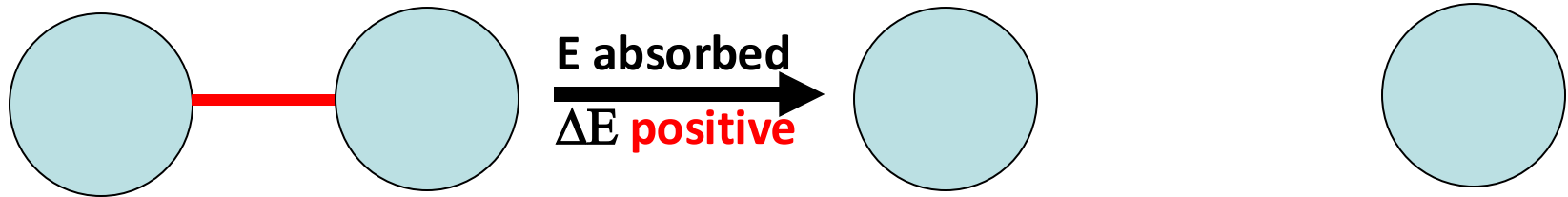
$$\begin{aligned} \text{D} \rightarrow \text{E}: q &= (m) \Delta H_{\text{vap}} \\ 10\text{g} (2444 \text{ J/g}) &= 24440 \text{ J} \end{aligned}$$

$$\begin{aligned} \text{E} \rightarrow \text{F}: q &= m (S_{\text{water vapor}}) \Delta T \\ 10\text{g} (2.00 \text{ J/gK}) (25 \text{ K}) &= 500 \text{ J} \end{aligned}$$

$$\begin{aligned} \text{TOTAL } 512.5 + 3340 \text{ J} + 4184 \text{ J} + 24440 \text{ J} &= \\ &= \mathbf{32977 \text{ J}} \end{aligned}$$

Why does anything produce or absorb energy?

The glue



Average Bond Enthalpies

Single Bonds

C—H	413	N—H	391	O—H	463	F—F	155
C—C	348	N—N	163	O—O	146		
C—N	293	N—O	201	O—F	190	Cl—F	253
C—O	358	N—F	272	O—Cl	203	Cl—Cl	242
C—F	485	N—Cl	200	O—I	234		
C—Cl	328	N—Br	243			Br—F	237
C—Br	276			S—H	339	Br—Cl	218
C—I	240	H—H	436	S—F	327	Br—Br	193
C—S	259	H—F	567	S—Cl	253		
		H—Cl	431	S—Br	218	I—Cl	208
Si—H	323	H—Br	366	S—S	266	I—Br	175
Si—Si	226	H—I	299			I—I	151
Si—C	301						
Si—O	368						
Si—Cl	464						

Multiple Bonds

C=C	614	N=N	418	O ₂	495
C≡C	839	N≡N	941		
C=N	615	N=O	607	S=O	523
C≡N	891			S=S	418
C=O	799				
C≡O	1072				

- Average bond enthalpies are positive, because bond **breaking** is an **endothermic** process.

Average Bond Enthalpies

NOTE: These are *average* bond enthalpies, not absolute bond enthalpies; the C—H bonds in methane, CH₄, will be a bit different than the C—H bond in chloroform, CHCl₃. This is a *less* accurate Measure of Enthalpy.

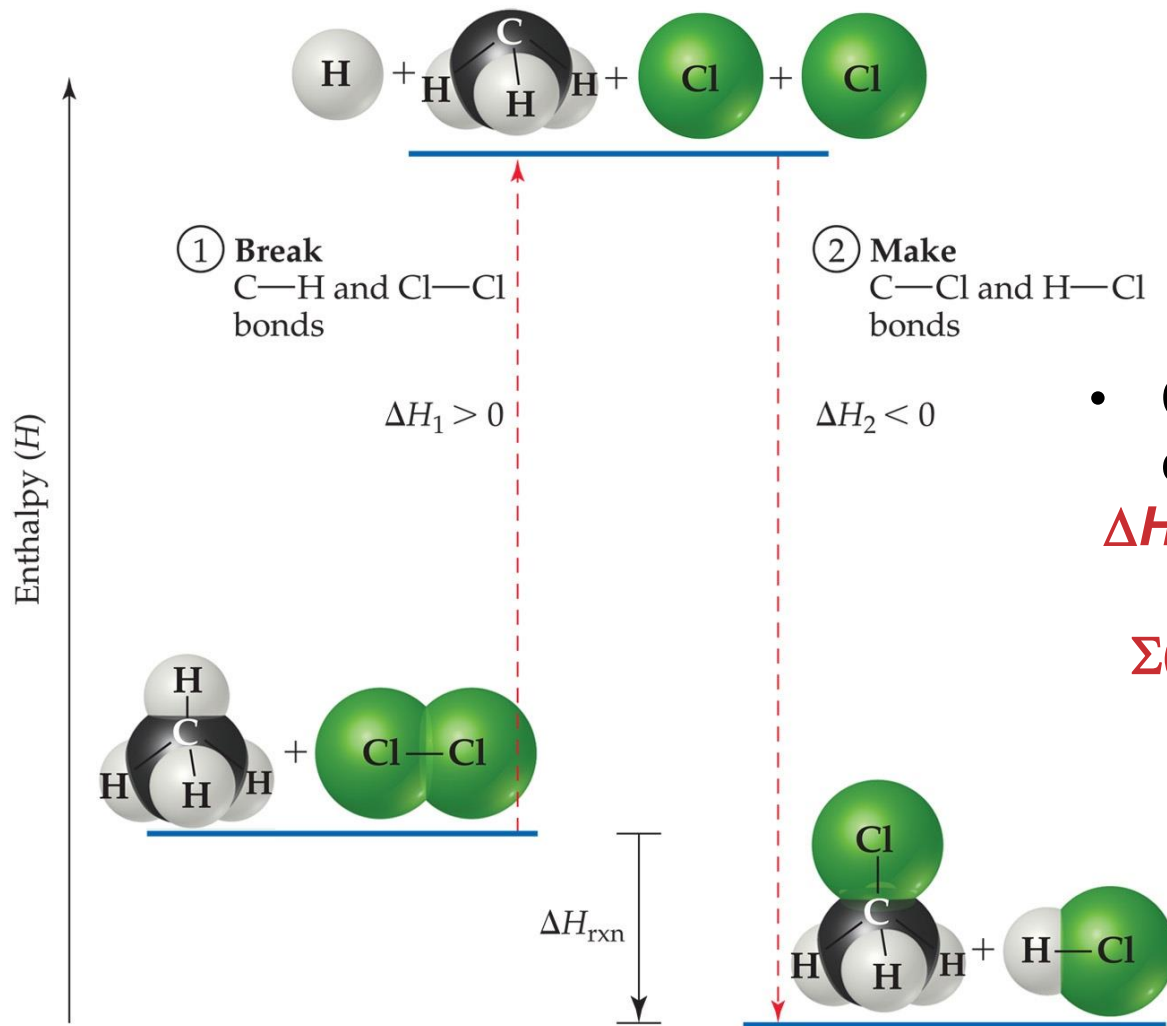
Single Bonds

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C—O	358	N—F	272	O—Cl	203	Br—F	237
C—F	485	N—Cl	200	O—I	234	Br—Cl	218
C—Cl	328	N—Br	243	S—H	339	Br—Br	193
C—Br	276	H—H	436	S—F	327	I—Cl	208
C—I	240	H—F	567	S—Cl	253	I—Br	175
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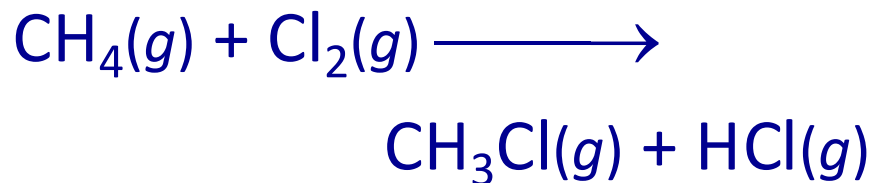
Enthalpies of Reaction



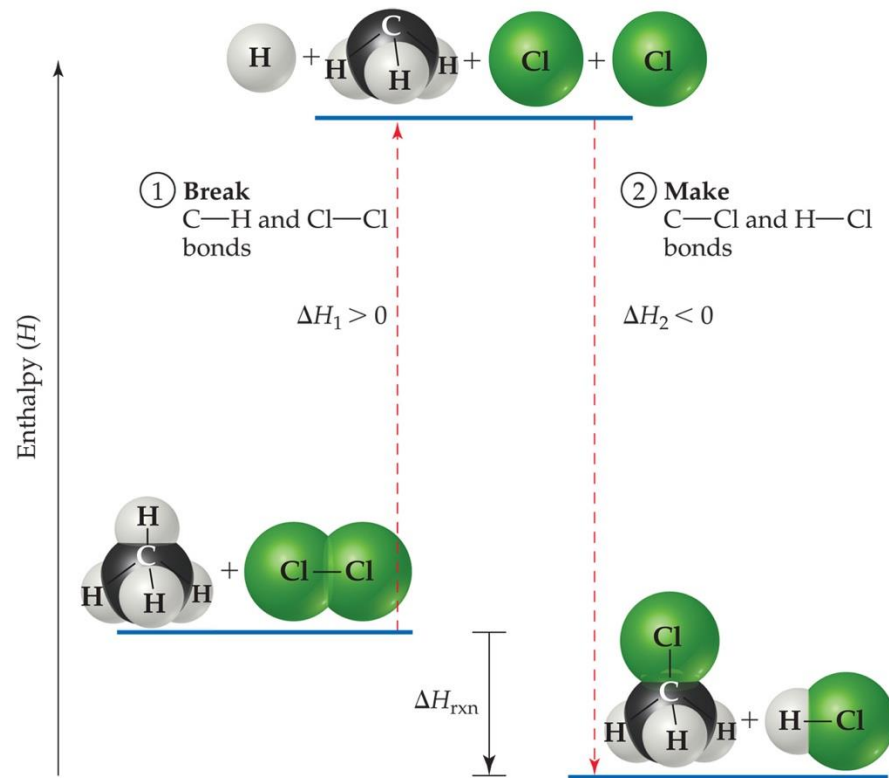
- Can use bond enthalpies to estimate ΔH for a reaction
$$\Delta H_{\text{rxn}} = \Sigma(\text{bond enthalpies of bonds broken}) - \Sigma(\text{bond enthalpies of bonds formed})$$

This is a fundamental idea in chemical reactions. **The heat of a reaction comes from breaking bonds and remaking bonds.**

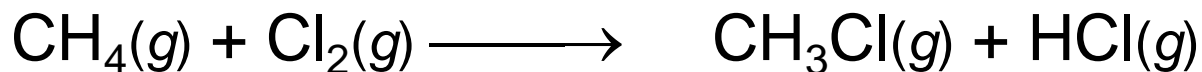
Enthalpies of Reaction



In this example, one C—H bond and one Cl—Cl bond are broken; one C—Cl and one H—Cl bond are formed.



Enthalpies of Reaction



So, Bonds broken - Bonds made

$$\begin{aligned}\Delta H_{\text{rxn}} &= [D(\text{C}-\text{H}) + D(\text{Cl}-\text{Cl})] - [D(\text{C}-\text{Cl}) + D(\text{H}-\text{Cl})] \\ &= [(413 \text{ kJ}) + (242 \text{ kJ})] - [(328 \text{ kJ}) + (431 \text{ kJ})] \\ &= (655 \text{ kJ}) - (759 \text{ kJ}) \\ &= -104 \text{ kJ}\end{aligned}$$

Enthalpies of Reaction

Which of the following is the most *exothermic* reaction?



Single Bonds

C—H	413	N—H	391	O—H	463	F—F	155
C—C	348	N—N	163	O—O	146		
C—N	293	N—O	201	O—F	190	Cl—F	253
C—O	358	N—F	272	O—Cl	203	Cl—Cl	242
C—F	485	N—Cl	200	O—I	234		
C—Cl	328	N—Br	243			Br—F	237
C—Br	276			S—H	339	Br—Cl	218
C—I	240	H—H	436	S—F	327	Br—Br	193
C—S	259	H—F	567	S—Cl	253		
		H—Cl	431	S—Br	218	I—Cl	208
Si—H	323	H—Br	366	S—S	266	I—Br	175
Si—Si	226	H—I	299			I—I	151
Si—C	301						
Si—O	368						
Si—Cl	464						

Multiple Bonds

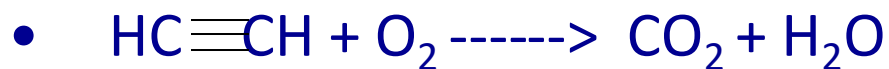
Bond Enthalpy and Bond Length

Bond	Bond Length (Å)	Bond	Bond Length (Å)
C—C	1.54	N—N	1.47
C=C	1.34	N=N	1.24
C≡C	1.20	N≡N	1.10
C—N	1.43	N—O	1.36
C=N	1.38	N=O	1.22
C≡N	1.16		
		O—O	1.48
C—O	1.43	O=O	1.21
C=O	1.23		
C≡O	1.13		

- We can also measure an average bond length for different bond types.
- As the number of bonds between two atoms increases, the bond length decreases.

Enthalpy problem:

- Calculate the enthalpy of reaction for:



Single Bonds

C—H	413	N—H	391	O—H	463	F—F	155
C—C	348	N—N	163	O—O	146	Cl—F	253
C—N	293	N—O	201	O—F	190	Cl—Cl	242
C—O	358	N—F	272	O—Cl	203	Br—F	237
C—F	485	N—Cl	200	O—I	234	Br—Cl	218
C—Cl	328	N—Br	243	S—H	339	Br—Br	193
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Si—Si	226	H—I	299				
Si—C	301						
Si—O	368						
Si—Cl	464						

Multiple Bonds

C=C	614	N=N	418	O ₂	495
C≡C	839	N≡N	941	S=O	523
C=N	615	N=O	607	S=S	418
C≡N	891				
C=O	799				
C≡O	1072				

Enthalpy problem:

- Calculate the enthalpy of reaction for:
- $\text{CH}_4 + 3/2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$
- $4(\text{C-H}) + 3/2(\text{O=O}) - 2(\text{C=O}) - 4(\text{OH})$
- $4(413) + 3/2(495) - 2(800) - 4(463) = -563 \text{ k}$

- $\text{HC}\equiv\text{CH} + 5/2\text{O}_2 \rightarrow 2\text{CO}_2 + \text{H}_2\text{O}$
- $1(\text{CC}) + 2\text{CH} + 5/2(\text{O=O}) - 4(\text{C=O}) - 2(\text{OH})$
- $1(834) + 2(413) + 5/2(495) - 4(800) - 2(463)$

Single Bonds

C—H	413	N—H	391	O—H	463
C—C	348	N—N	163	O—O	146
C—N	293	N—O	201	O—F	190
C—O	358	N—F	272	O—Cl	203
C—F	485	N—Cl	200	O—I	234
C—Cl	328	N—Br	243	S—H	339
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Energy in Foods

Most of the fuel in the food we eat comes from carbohydrates and fats.

TABLE 5.4 Compositions and Fuel Values of Some Common Foods

	Approximate Composition (% by mass)			Fuel Value	
	Carbohydrate	Fat	Protein	kJ/g	kcal/g (Cal/g)
Carbohydrate	100	–	–	17	4
Fat	–	100	–	38	9
Protein	–	–	100	17	4
Apples	13	0.5	0.4	2.5	0.59
Beer*	1.2	–	0.3	1.8	0.42
Bread	52	3	9	12	2.8
Cheese	4	37	28	20	4.7
Eggs	0.7	10	13	6.0	1.4
Fudge	81	11	2	18	4.4
Green beans	7.0	–	1.9	1.5	0.38
Hamburger	–	30	22	15	3.6
Milk (whole)	5.0	4.0	3.3	3.0	0.74
Peanuts	22	39	26	23	5.5

*Beers typically contain 3.5% ethanol, which has fuel value.

What's the deal with fat?

Note: Your life is one, long slow combustion reaction

- Carbohydrates:



- Fats: more steps

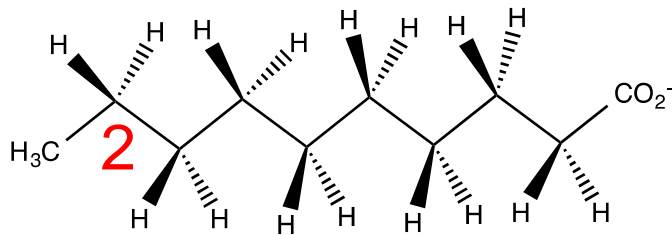
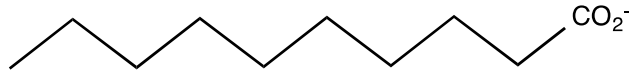


Fat storage.

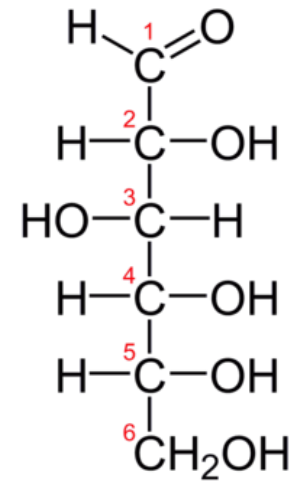
It also clogs your arteries.

Energy and oxidation states

- Oxidation state of C in a fatty acid:



- Oxidation state of C in glucose:

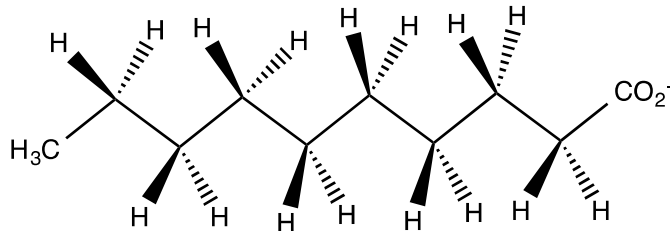
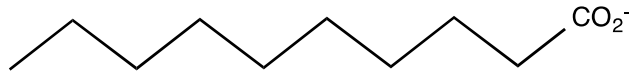


What's the oxidation number for C2
In the fatty acid above?

- A. +4
- B. +2
- C. 0
- D. -2
- E. -4

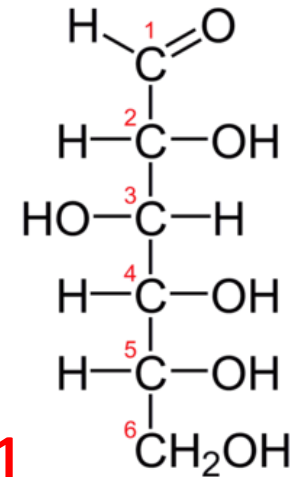
Energy and oxidation states

- Oxidation state of C in a fatty acid:



-3 -2 -2 -2 -2 -2 -2 -2 +3

- Oxidation state of C in glucose:



- $\text{C}_6\text{H}_{12}\text{O}_6$

• -1,0,0,0,0,+1

- Both go to CO_2 C: +4

REview

- Which do you think has the most negative heat of combustion?
- A. CH_4
- B. CH_3OH

REview

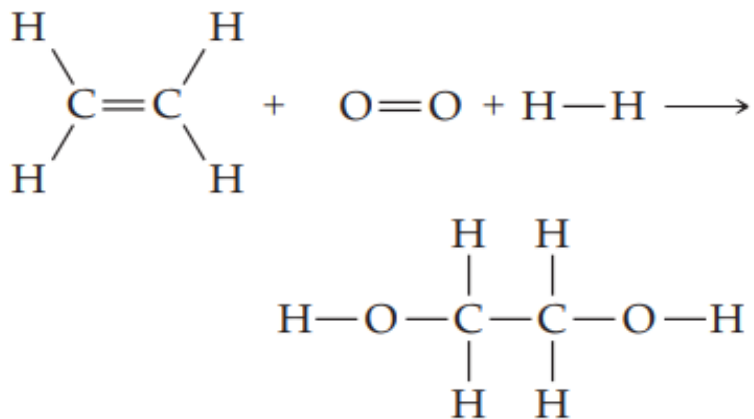
- Which do you think has the most negative heat of combustion?
 - A. CH_4 . -891
 - B. CH_3OH . -727
- Given this which bond do you think is stronger?
 - A. O-H
 - B. C-H

Average bond enthalpies of some bonds are given in the table below.

Bond	Average Bond Enthalpies (kJ/mol)
C–C	348
C=C	614
C–H	413
O=O	495
H–H	436
H–O	463
C–O	358

Bonds broken – bonds made

$$1(\text{C}=\text{C}) + 1\text{O}=\text{O} + 1\text{H}-\text{H} - 2(\text{C}-\text{O}) - 2(\text{O}-\text{H}) - 1\text{C}-\text{C} = 614 + 495 + 436 - 2(358) - 2(413) - 348 = -445 \text{ kJ}.$$

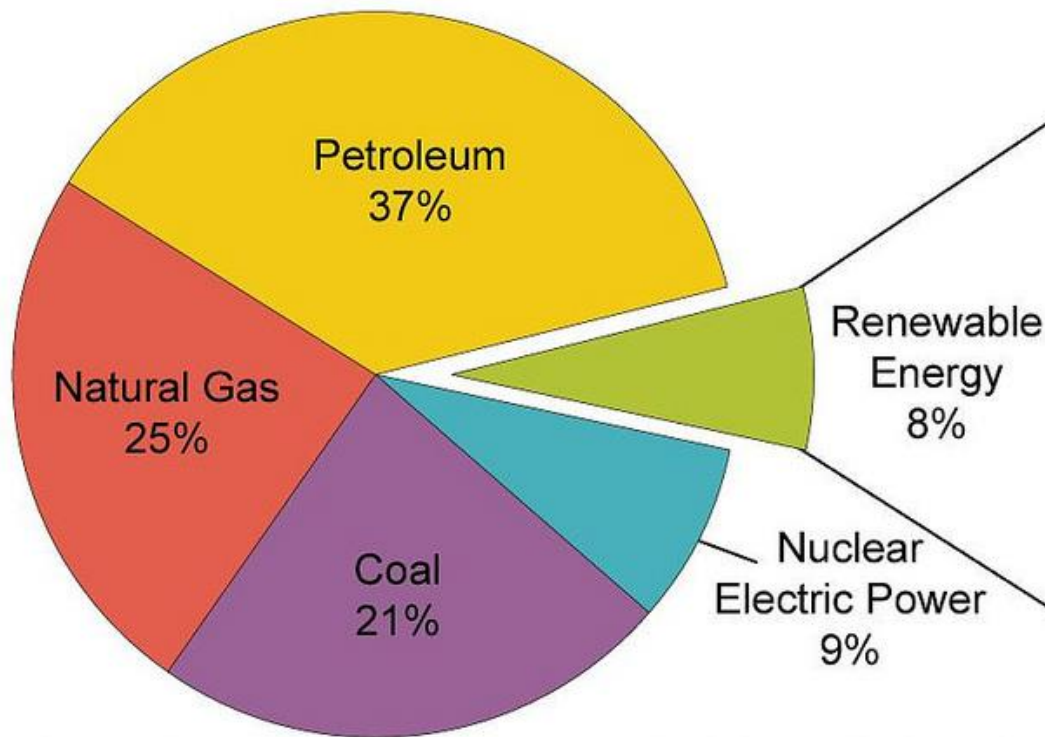


Fuels

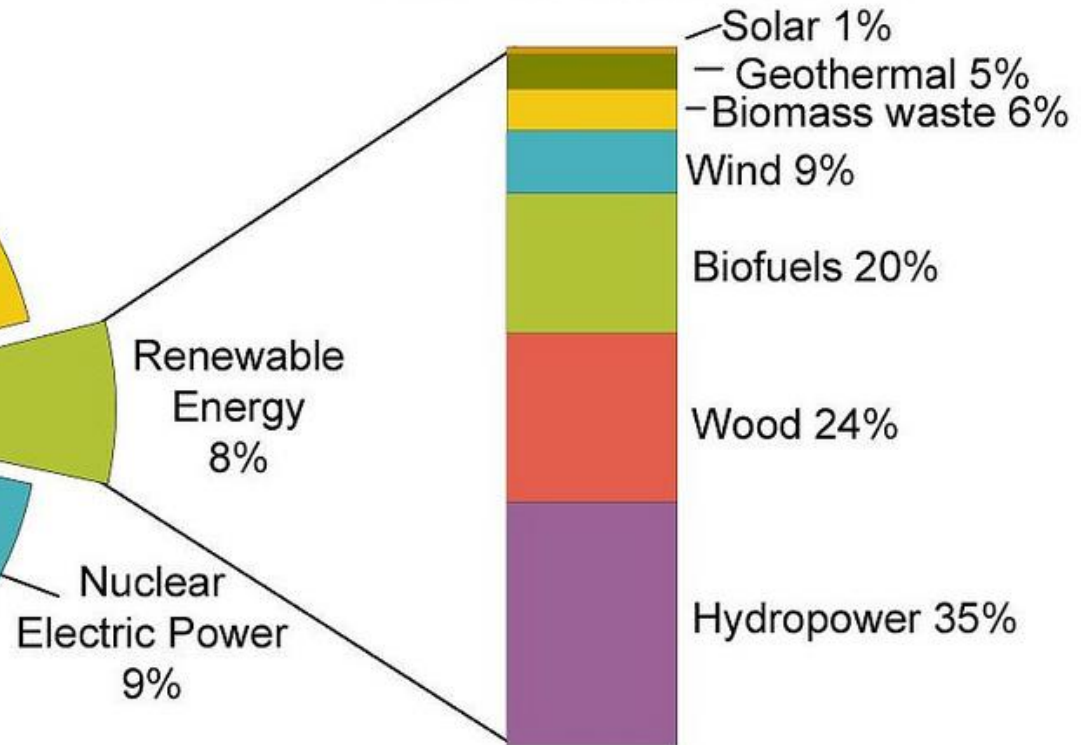
The vast majority of the energy consumed in this country comes from fossil fuels.

U.S. Energy Consumption by Energy Source, 2009

Total = 94.578 Quadrillion Btu



Total = 7.744 Quadrillion Btu



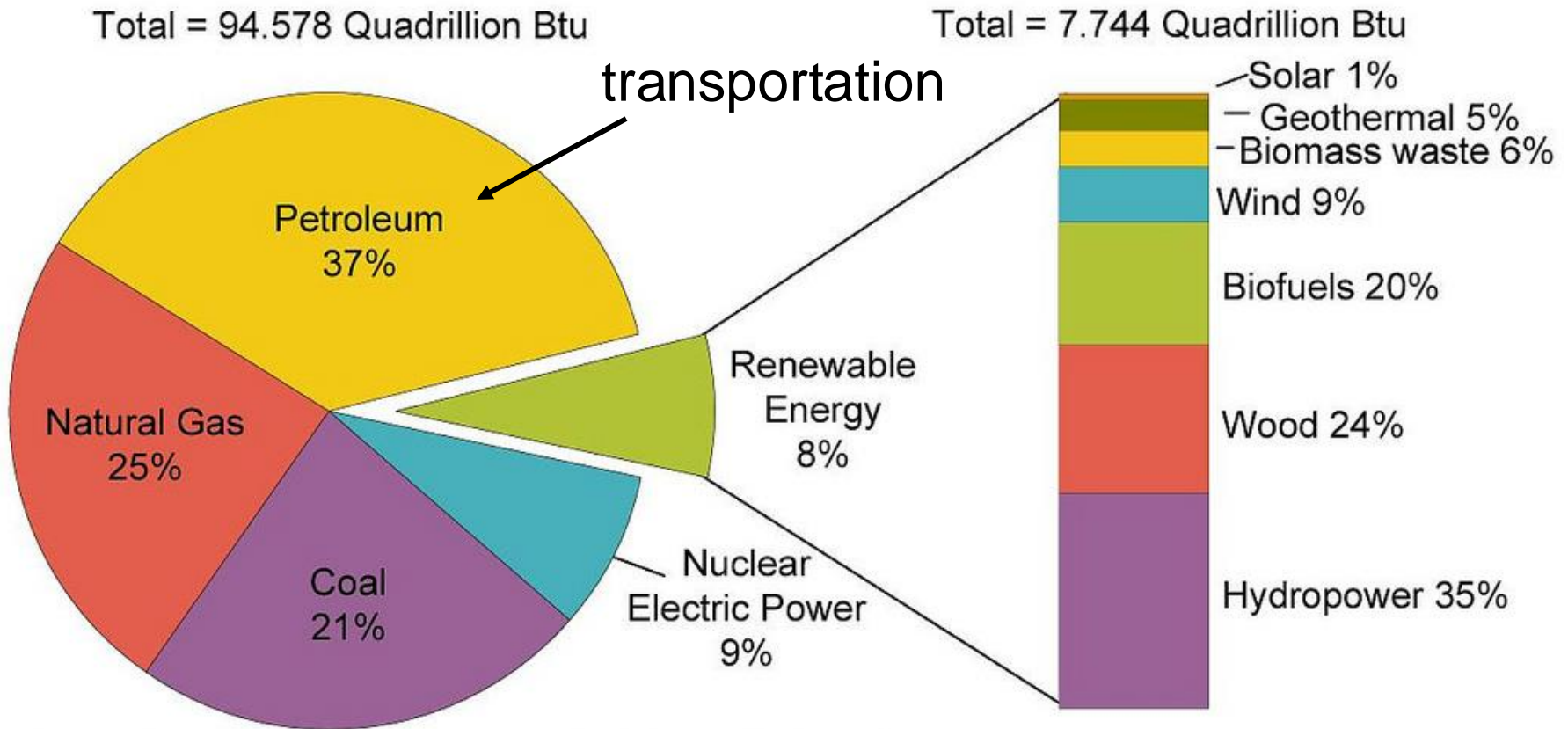
Note: Sum of components may not equal 100% due to independent rounding.

Source: U.S. Energy Information Administration, *Annual Energy Review 2009*, Table 1.3, Primary Energy Consumption by Energy Source, 1949-2009 (August 2010).

Major issues

- Portable fuel (liquid, relatively light), transportation
- Non-portable fuel (makes electricity).

U.S. Energy Consumption by Energy Source, 2009



Note: Sum of components may not equal 100% due to independent rounding.

Source: U.S. Energy Information Administration, *Annual Energy Review 2009*, Table 1.3, Primary Energy Consumption by Energy Source, 1949-2009 (August 2010).

The Energy cycle:

- Us (and almost everything else alive on the earth):
- $C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$
- Fossil fuel production:
- $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ (for methane)
- Plants:
- $6CO_2 + 6H_2O + \text{light} \rightarrow C_6H_{12}O_6 + 6O_2$

Net CO_2 production could therefore be 0.

Energy research on Campus

ADREC

Anaerobic digester



Energy research on Campus

ADREC

Anaerobic digester

- Anaerobic Digestion:

bacteria



Basically:



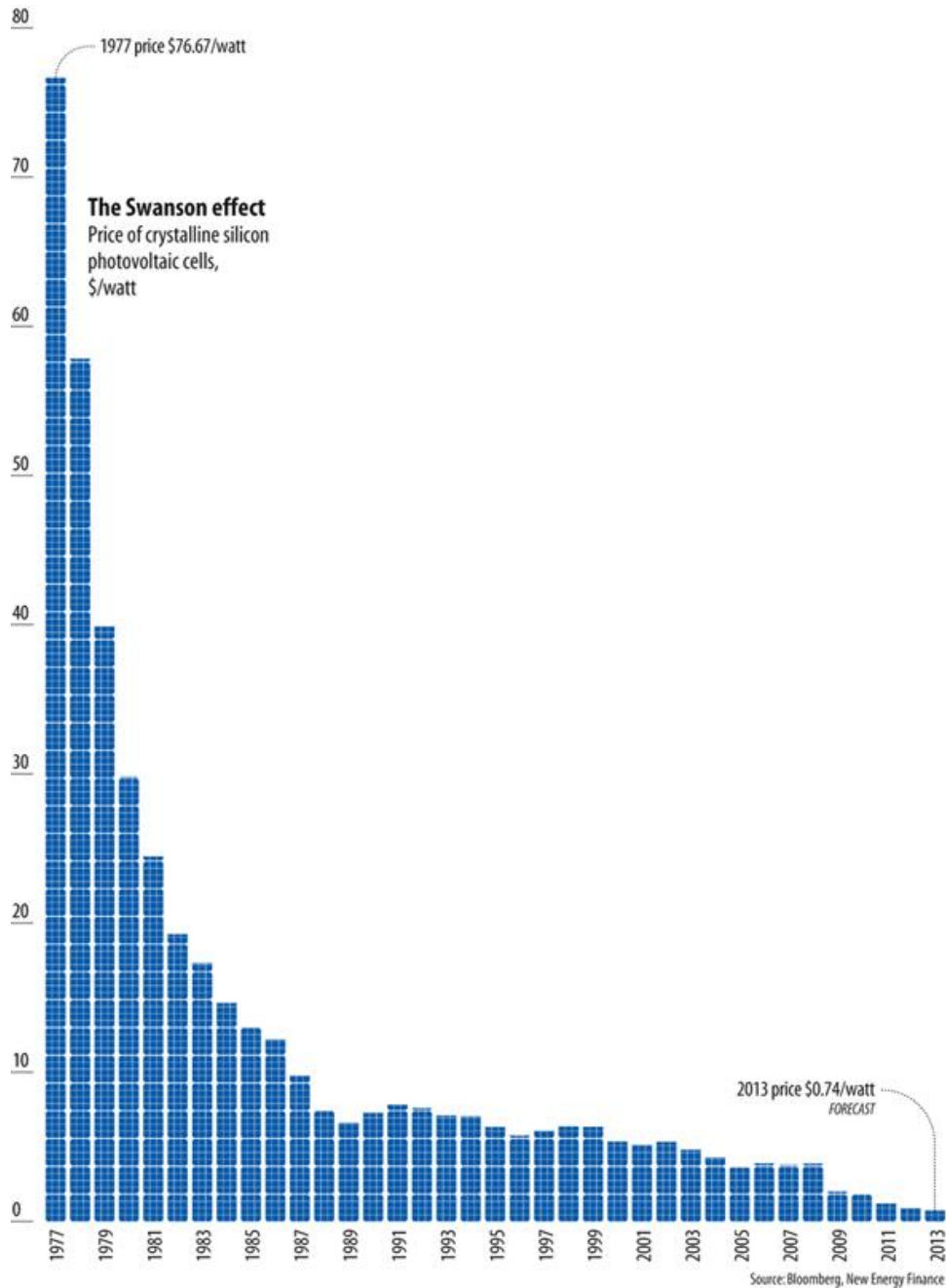
<https://www.egr.msu.edu/bae/adrec/feature/south-campus-anaerobic-digester>

The problem with oil

- Not “renewable” (will run out)
- Pollution (combustion not perfect).
- Global warming
CO₂ absorbs heat.



The end of fuel?



Very soon, only idiots will not have solar panels on their houses. It'll just be too cheap not to.

Battery prices are also plummeting

- <https://electrek.co/2017/01/30/electric-vehicle-battery-cost-dropped-80-6-years-227k>

The end of fuel?

<https://www.wired.com/story/general-motors-electric-cars-plan-gm/>

The Nobel Prize in Chemistry 2020

Jennifer Doudna,

Emmanuelle Charpentier

CRISPR

What is CRISPR?



The Nobel Prize in Chemistry 2020

Jenifer Doudna,

Emmanuelle Charpentier

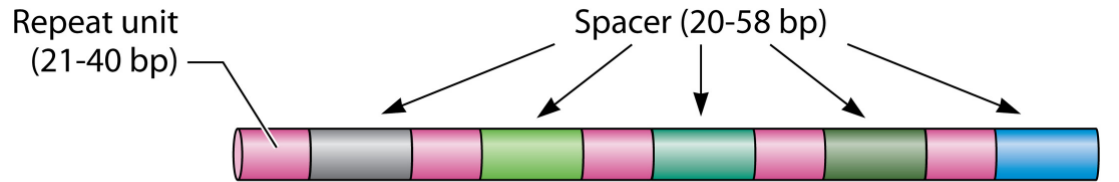
CRISPR

What is CRISPR?



- CRISPR:
 - **Clustered, Regularly Interspaced Palindromic Repeats.**
 - Originally observed as a weird DNA sequence in bacteria. Who cares? So what?

Observation 1 1987



E. coli (1987)

CGGTTTATCCCCGCTGCGCGGGGAACTC

H. mediterranei (1993)

GTTACAGACGAACCCTAGTTGGGTTGAAGC



G
T C
C•G
G•C
C•G
C•G
C•G
C•G
C•G
T•A

GCCTTT^A AACTC



G
A•T
T T
C•G
C•G
C•G
A•T
A•T

GTTACAGACG GAAGC

- **CRISPR:**

- **Clustered, Regularly Interspaced Palindromic Repeats. Data, just DNA sequence.**

The mechanism of adapti

CAS1 and 2.

Invading DNA recognized by Cas proteins

Chopped up and stuck into spacer regions

Now stored in genome passed on

1 Repeat and spacer DNA transcribed
(RNA made from DNA)

2. Chopped up into units of
1 repeat, one spacer (CrRNA)

3. Cas9 binds CrRNA. Chops DNA
That matches that sequence.
Kill invading viruses DNA.

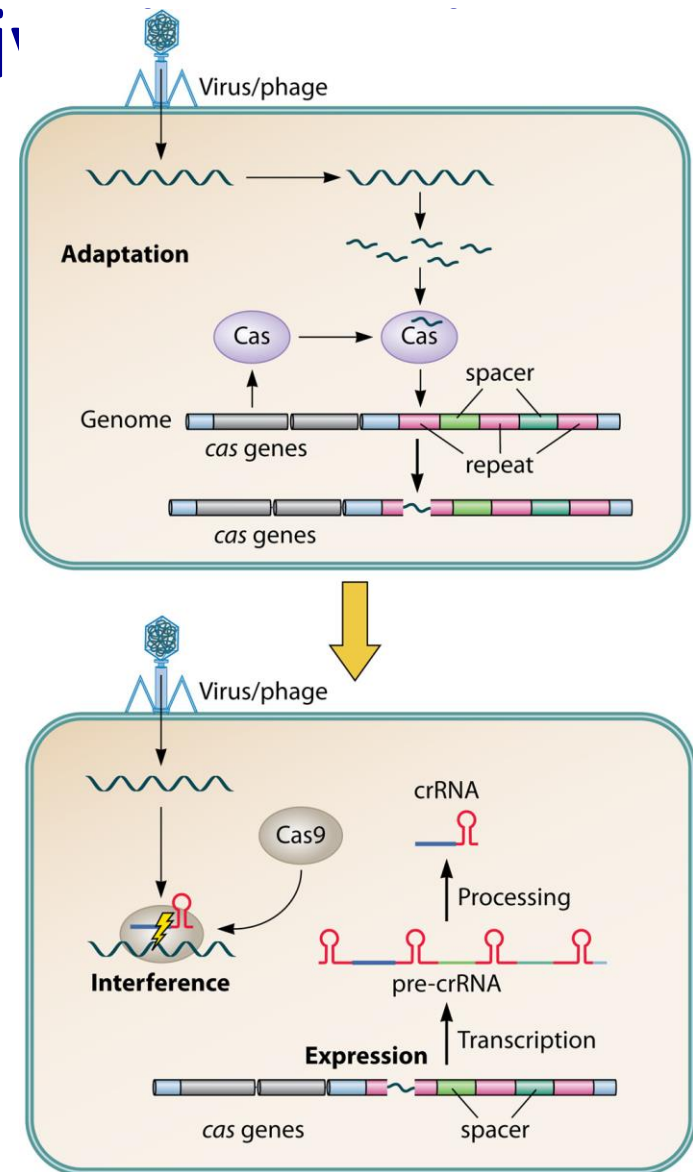


FIG 4 Process of CRISPR-Cas acquired immune system. (Top) Adaptation. The invading DNA is recognized by Cas proteins, fragmented and incorporated into the spacer region of CRISPR, and stored in the genome. Expression (bottom). Pre-crRNA is generated by transcription of the CRISPR region and is processed into smaller units of RNA, named crRNA. (Bottom) Interference. By taking advantage of the homology of the spacer sequence present in crRNA, foreign DNA is captured, and a complex with Cas protein having nuclease activity cleaves DNA.

Gene Editing

If you can specifically cut a genome
in a living cell....

You can edit it anyway you want.

If you do it to the germ line (egg, sperm)

You change the organism forever.

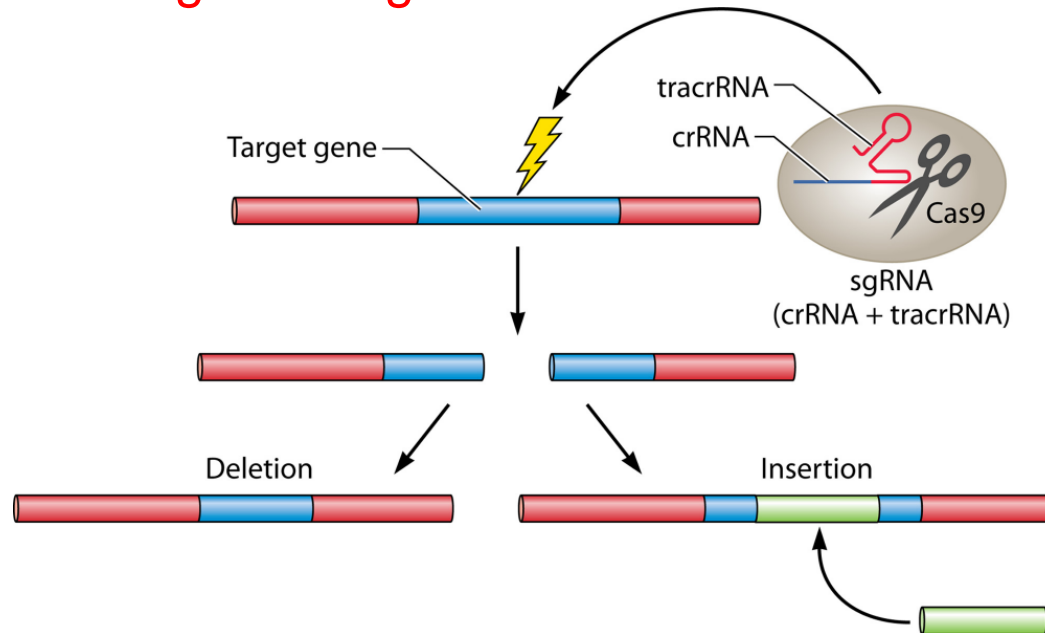
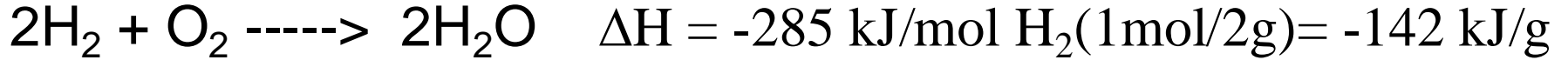


FIG 5 Genome editing by CRISPR-Cas9. The principle of genome editing is the cleavage of double-stranded DNA at a targeted position on the genome. The type II is the simplest as a targeted nuclease among the CRISPR-Cas systems. The CRISPR RNA (crRNA), having a sequence homologous to the target site, and *trans*-activating CRISPR RNA (tracrRNA) are enough to bring the Cas9 nuclease to the target site. The artificial linkage of crRNA and tracrRNA into one RNA chain (single-guide RNA [sgRNA]) has no effect on function. Once the Cas9-sgRNA complex cleaves the target gene, it is easy to disrupt the function of the gene by a deletion or insertion mutation. This overwhelmingly simple method is now rapidly spreading as a practical genomic editing technique.

Hydrogen, the perfect fuel?



This is literally what fuel cells do. You get nothing but water!

Efficiency/conservation

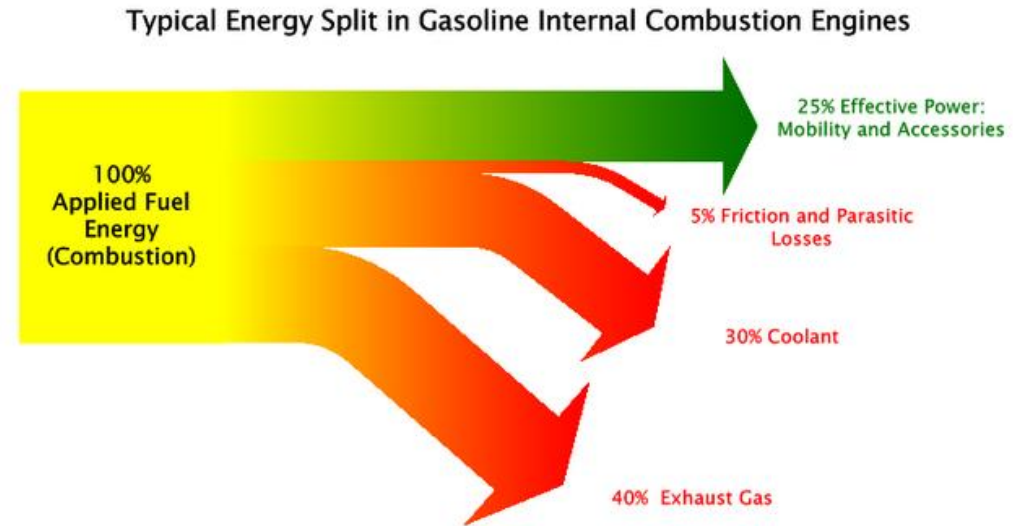
- U.S. could decrease energy needs by 20-50% by being less wasteful.
- High mileage cars
- more energy efficient building/homes.

Hybrid car

- Gas engine plus electric motor
- Why?
- All the energy is still coming from burning gasoline.

Hybrids

- Electric motors are way more efficient than gas engines. (94%)
- Note, your engine is very hot,
- It must be cooled
- Flush all that E down drain. No work, only heat.

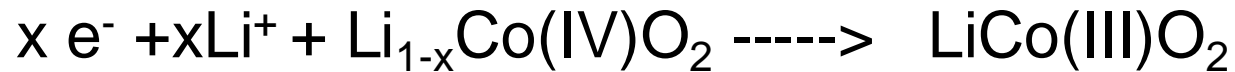


gas engines are 24-30% efficient

Problem: batteries suck!

Heavy, expensive, limited recharging cycles, limited current etc.

Li ion battery



Lithium is really light.

Dissolves in organic solvents which are also light.

Li is at the top of the activity series. Means a higher potential (more voltage per battery cell)

Hybrids

- Electric motors work at low speeds
- gas engine shuts off when not needed
- at low speeds, stop lights, etc.
- (infinite torque, really go from 0-15)
- Gas engine charges battery and is used at higher speeds
- Hybrids get BETTER gas milage in town versus highway

Other sources

How much bang for your buck?

	Approximate Elemental Composition (mass %)			Fuel Value (kJ/g)
	C	H	O	
Wood (pine)	50	6	44	18
Anthracite coal (Pennsylvania)	82	1	2	31
Bituminous coal (Pennsylvania)	77	5	7	32
Charcoal	100	0	0	34
Crude oil (Texas)	85	12	0	45
Gasoline	85	15	0	48
Natural gas	70	23	0	49
Hydrogen	0	100	0	142

The problem with Hydrogen

Storage

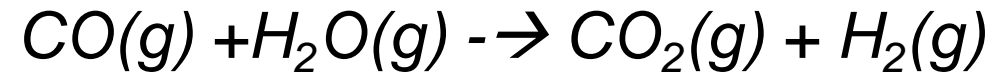
gas, less dense, hard to get enough in the car and have trunk space

Kaboom (Hindenburg)

Where do you get the hydrogen?

The problem with Hydrogen

Where do you get the hydrogen? (*petroleum*)



Ethanol, where does it come from

- Alcoholic fermentation:
- $C_6H_{12}O_6 \rightarrow 2CO_2 + 2C_2H_5OH$ (ethanol) $\Delta H = -76 \text{ kJ/mol}$
- $-1270 \qquad 2(-393) \qquad 2(-280)$
- (anaerobic, bacteria & yeast can do this, we can't)

Exactly the same place it comes from in your beer.

Ethanol

- Alcoholic fermentation:
- $C_6H_{12}O_6 \xrightarrow{\text{bug}} 2CO_2 + 2C_2H_6O$ (ethanol) $\Delta H = -76 \text{ kJ/mol}$
- $-1270 \quad 2(-393) \quad 2(-280)$
- (anaerobic, yeast can do this, we can't) only to 10%.
- Distillation (requires energy) to purify.

Alcohol combustion:



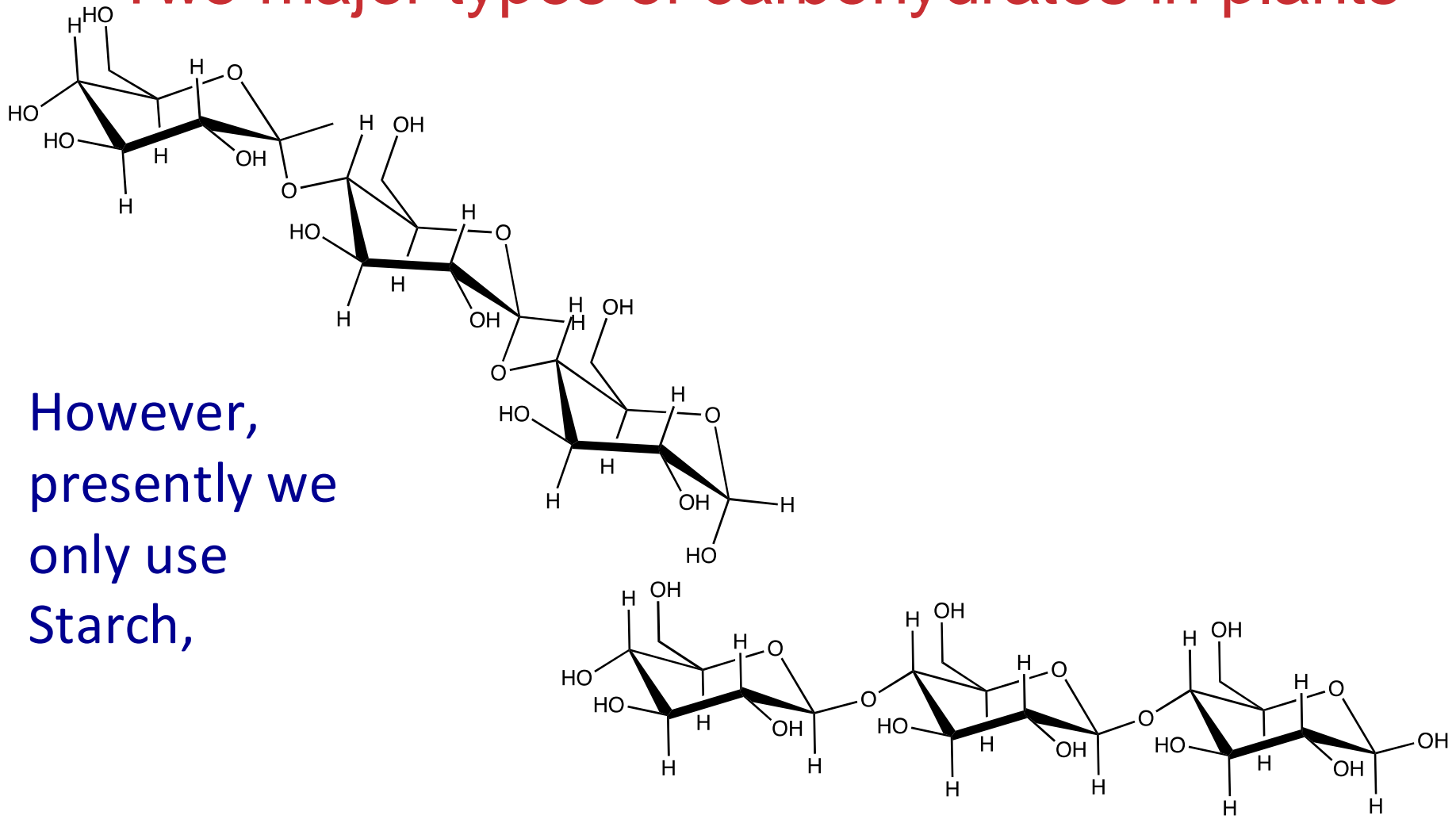
But why would this be better for global warming?

Ethanol, problems

- Lots of land to grow (yield 2-4 tons/acre)
- All present agricultural land in U.S. would not be enough for all transportation needs.
- requires fertilizer, tractors, etc. for growing (energy)
- Distillation requires energy
- For every 1.4 kJ need 1.0 kJ, much more than oil
- Brazil, however, is approaching 50% ethanol for transportation
- Why? Sugar cane, largest starch or sugar yield/acre.
- But, you can't grow sugar cane on the great plains.

Ethanol

Two major types of carbohydrates in plants



- However, presently we only use Starch,

not cellulose

Most stuff in plants is cellulose

Cellulosic ethanol

- 10+ tons/acre (as opposed to 2-4 tons/acre)
- Can use any crop, not just food crops with high starch (“switch grass”).
- Problem: Breaking it down to small sugars that yeast can ferment.
- Need cellulase, the enzyme that breaks this up.
- This is a comparatively easy problem to solve
- (compared to hydrogen.)

Ethanol can work.

Things to consider

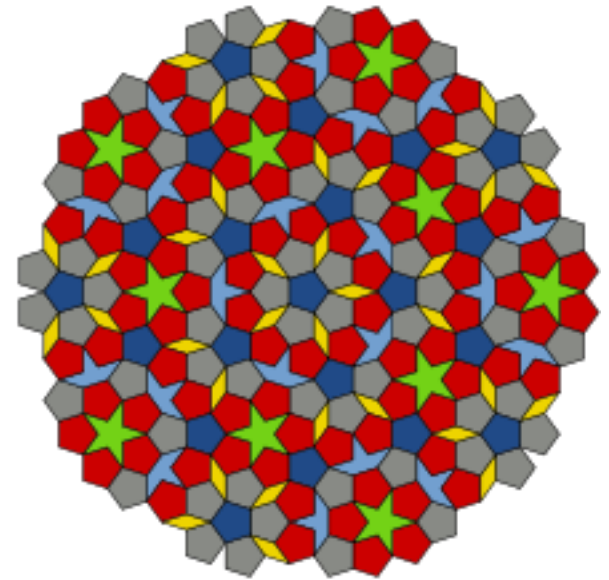
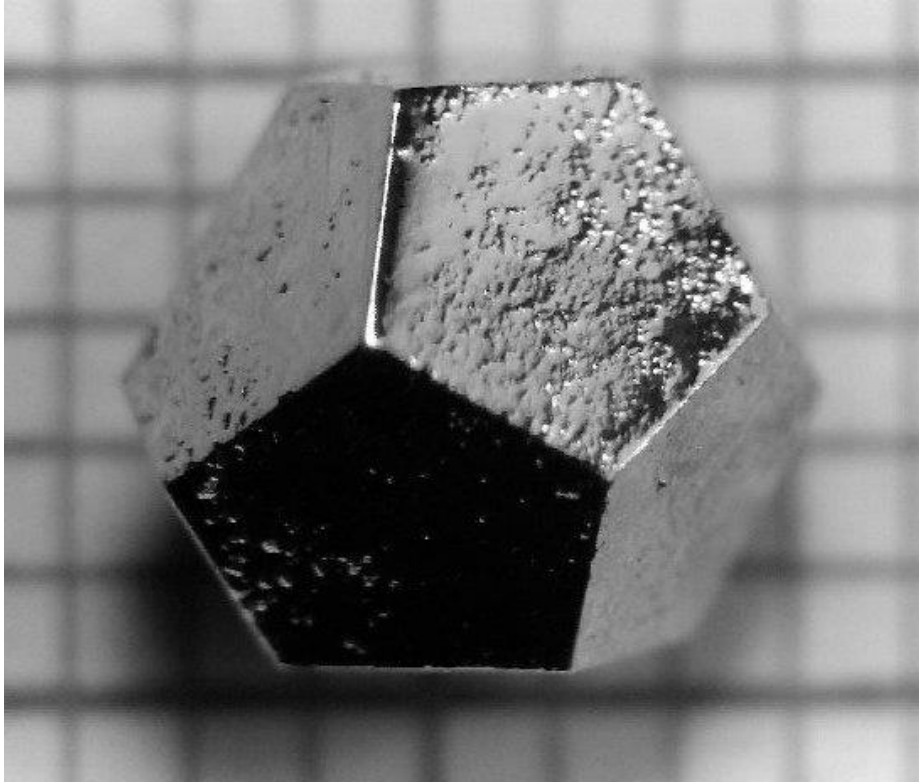
- Energy yield (how much E out versus E in)?
- Break even price (how much/gallon of gas equivalents (present corn ethanol is 2.25/gallon just to make)).
- Where is the technology NOW?
- Is storage required, & if so, how you gonna do it
- (solar when the sun doesn' t shine)
- Remember, at present Batteries suck!

The Chemistry Nobel Prize

- Daniel Shechtman,
- Technion, Israel
- For:
- The discovery of “quasi-crystals” in 1984

The Chemistry Nobel Prize

- An Ho-Mg-Zn quasi-crystal



Note, the five-fold symmetry of the faces!
This was thought to be impossible!
Is this a solid?

The Thermite reaction

- $2\text{Al} + \text{Fe}_2\text{O}_3 \text{ -----} \rightarrow \text{Al}_2\text{O}_3 + 2\text{Fe}$
- What kind of reaction is this?
- Why does it happen?
- Used for welding railroad tracks
- What is the heat of reaction given:
- $2\text{Fe} + 3/2\text{O}_2 \text{ -----} \rightarrow \text{Fe}_2\text{O}_3 \quad \Delta\text{H} = -825.5 \text{ KJ}$
- $2\text{Al} + 3/2\text{O}_2 \text{ -----} \rightarrow \text{Al}_2\text{O}_3 \quad \Delta\text{H} = -1675.7 \text{ KJ}$

The Thermite Reaction

- $2\text{Al} + \text{Fe}_2\text{O}_3 \longrightarrow \text{Al}_2\text{O}_3 + 2\text{Fe}$
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 - $2\text{Al} + 3/2\text{O}_2 \longrightarrow \text{Al}_2\text{O}_3 \quad \Delta H = -1675.7 \text{ KJ}$
 - $\text{Fe}_2\text{O}_3 \longrightarrow 2\text{Fe} + 3/2\text{O}_2 \quad \Delta H = 825.5 \text{ KJ}$
-
- $2\text{Al} + \text{Fe}_2\text{O}_3 \longrightarrow \text{Al}_2\text{O}_3 + 2\text{Fe} \quad \Delta H = -850.2 \text{ KJ}$

A thermite mystery:

<http://www.youtube.com/watch?v=BnHR4cMXiyM>