Chapter 5
Thermochemistry

The energy of chemical reactions.
How do you keep track of it?
Where does it come from?
Energy

• The ability to:
  • do work
  • transfer heat.

  ➢ Work: Energy used to cause an object that has mass to move.
  ➢ Heat: Energy used to cause the temperature of an object to rise.

We will see that both are due to molecular motion.
Potential Energy

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

More potential E

Less P.E. as bike goes down.
Kinetic Energy

Energy an object possesses by virtue of its motion.

\[ KE = \frac{1}{2} mv^2 \]
Units of Energy

• The SI unit of energy is the joule (J).
  \[ 1 \text{ J} = 1 \frac{\text{kg m}^2}{\text{s}^2} \]

• An older, non-SI unit is still in widespread use: The calorie (cal).
  \[ 1 \text{ cal} = 4.184 \text{ J} \]
Energy accounting

• Must identify where different types of Energy go.

• Therefore, must identify the places.
System and Surroundings

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

- Energy used to move an object over some distance.
- \( w = F \cdot d \),
- \( w \) = work,
- \( F \) = force
- \( d \) = distance over which the force is exerted.
Heat

- Energy can also be transferred as heat.
- Heat flows from warmer objects to cooler objects.
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. (1/2mv^2)
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. (1/2mv^2)

Ball hits ground, K.E. =0, but E has to go somewhere. So
1. Ball gets squashed
2. Heat comes out.
First Law of Thermodynamics

• Energy is conserved.
• In other words, the total energy of the universe is a constant;

\[ \Delta E_{\text{System}} = -\Delta E_{\text{surroundings}} \]
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 \( E \).

\[
E_{\text{internal, Total}} = E_{\text{K.E.}} + E_{\text{P.E.}} + E_{\text{electrons}} + E_{\text{nuclei}} + \ldots \ldots
\]

almost impossible to calculate total internal energy.

Instead, we are always looking at change in energy (\( \Delta E \))
Internal Energy

By definition, the change in internal energy, $\Delta E$, is the final energy of the system minus the initial energy of the system:

$$\Delta E = E_{\text{final}} - E_{\text{initial}}$$
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.
Changes in Internal Energy

- 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, $\Delta E = q + w$. 
$\Delta E, \ q, \ w, \ and \ Their \ Signs$

For $q$  
+ means system gains heat  
− means system loses heat

For $w$  
+ means work done on system  
− means work done by system

For $\Delta E$  
+ means net gain of energy by system  
− means net loss of energy by system

Surroundings suck heat out of water.

50 g $H_2O \ (l)$ 100°C  
- $q$  
Cooling

50 g $H_2O \ (l)$ 25°C  
+ $q$  
Heating

50 g $H_2O \ (s)$ 0°C  
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

- When heat is absorbed by the system from the surroundings, the process is **endothermic**.
Exchange of Heat between System and Surroundings

- When heat is absorbed by the system from the surroundings, the process is **endothermic**.
- When heat is released by the system to the surroundings, the process is **exothermic**.

\[ \Delta H < 0 \]  
(Exothermic)
State Functions

Total internal energy of a system:
K.E. + $E_{\text{electrons}} + E_{\text{nucleus}} + P.E_{\text{total}}$

virtually impossible to measure/calculate
State Functions

- However, we do know that 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 it’s PATH INDEPENDENT
- And so, $\Delta E$ depends only on $E_{\text{initial}}$ and $E_{\text{final}}$. 
State Functions

- However, $q$ and $w$ are not state functions.
- Whether the battery is shorted out or is discharged by running the fan, its $\Delta E$ is the same.
  - But $q$ and $w$ are different in the two cases.
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
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?
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 $\Delta E$ 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. both $q$ and $w$ not 0

b.2. $w$ 0, $q$ larger

c.1. $\Delta E$ the same in each case
Work

Now we can measure the work:

\[ w = -P \Delta V \]

\[ \text{Zn} + 2\text{HCl} \quad \rightarrow \quad \text{H}_2(\text{g}) + \text{ZnCl}_2 \]
Enthalpy($H$)

\[ H = E + PV \]

This is the definition of Enthalpy for any process. But why do we care?
Enthalpy

\[ H = E + PV \]

- at constant pressure, \( \Delta H \), is

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

- This can be written (if \( P \) constant)

\( \Delta H = \Delta E + P\Delta V \)
Enthalpy

• Since $\Delta E = q + w$ and $w = -P\Delta V$
  substitute these into the enthalpy expression:

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

$$\Delta H = (q+w) - w$$

$$\Delta H = q$$

• Note: true at constant pressure

• $q$ is a state function at const P & only PV work.
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

- A process is endothermic when $\Delta H$ is positive.
Endothermicity and Exothermicity

- A process is endothermic when $\Delta H$ is positive.
- A process is exothermic when $\Delta H$ is negative.
Thermochemistry

Enthalpies of Reaction

The *change* in enthalpy, $\Delta H$, is the enthalpy of the products minus the enthalpy of the reactants:

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

For the reaction $\text{CH}_4(g) + 2 \text{O}_2(g) \rightarrow \text{CO}_2(g) + 2 \text{H}_2\text{O}(l)$,

$$\Delta H_1 = -890 \text{ kJ}$$

$$\Delta H_2 = 890 \text{ kJ}$$
Enthalpies of Reaction

This quantity, $\Delta H$, is called the enthalpy of reaction, or the heat of reaction.
The Truth about Enthalpy

1. Enthalpy is an extensive property.
2. $\Delta H$ for a reaction in the forward direction is equal in size, but opposite in sign, to $\Delta H$ for the reverse reaction.
3. $\Delta H$ for a reaction depends on the state of the products and the state of the reactants.
Enthalpy of reaction example

Consider the reaction:

$$2\text{KClO}_3 \rightarrow 2\text{KCl} + 3\text{O}_2 \quad \Delta H = -89.4 \text{ kJ/mol}$$

a. What is the enthalpy change for formation of 0.85 moles of $\text{O}_2$?
Enthalpy of reaction example

Consider the reaction:

\[ 2 \text{KClO}_3 \rightarrow 2 \text{KCl} + 3 \text{O}_2 \quad \Delta H = -89.4 \text{ kJ/mol} \]

a. What is the enthalpy change for formation of 0.85 moles of \( \text{O}_2 \)?

\[ 2 \text{KClO}_3 \rightarrow 2 \text{KCl} + 3 \text{O}_2 \quad \Delta H = -89.4 \text{ kJ/mol} \]

\[ 0.855 \text{ mol} \]

\[ \Delta H = -89.4 \text{ kJ/3 mol } \text{O}_2(0.855 \text{ mol } \text{O}_2) = -25.5 \text{ kJ} \]
Since we cannot know the exact enthalpy of the reactants and products, we measure $\Delta H$ through calorimetry, the measurement of heat flow.
Heat Capacity and Specific Heat

• The amount of energy required to raise the temperature of a substance by 1 K (1 °C) is its **heat capacity**.

• We define **specific heat capacity** (or simply **specific heat**) as the amount of energy required to raise the temperature of 1 g of a substance by 1 K.
Heat Capacity and Specific Heat

Specific heat, then, is

\[ s = \frac{q}{m \times \Delta T} \]
Constant Pressure Calorimetry

By carrying out a reaction in aqueous solution in a simple calorimeter such as this one, one can indirectly measure the heat change for the system by measuring the heat change for the water in the calorimeter.
Because the specific heat for water is well known (4.184 J/mol-K), we can measure $\Delta H$ for the reaction with this equation:

$$q = m \times s \times \Delta T$$
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 $\Delta H$ (in kJ/mol ammonium nitrate) for the solution process. Assume that the specific heat is constant and $= 1.0 \text{ g/ml}^\circ\text{C}$. (b) Is this process endothermic or exothermic?
Thermochemistry

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 $\Delta H$ (in kJ/mol ammonium nitrate) for the solution process. Assume that the specific heat is constant and $s = 1.0 \text{ g/ml}^\circ\text{C}$. (b) Is this process endothermic or exothermic?

Reaction:
\[ \text{NH}_4\text{NO}_3(s) \rightarrow \text{NH}_4^+(aq) + \text{NO}_3^-(aq) \]

\begin{align*}
gr &\quad 3.88 \text{ g} \\
MW &\quad 80.04 \text{ g/mol} \\
Mol &\quad 0.0484 \text{ mol} \\
\end{align*}

\[ q = s(specific \ heat)m(mass)\Delta T \]
\[ q = s(J/\text{g}^\circ\text{C})m(grams)(T_{\text{final}} - T_{\text{initial}}) \]
\[ q_{\text{water}} = 4.184(J/\text{g}^\circ\text{C})(60.0 \text{ g})(18.4^\circ\text{C} - 23.0^\circ\text{C}) = -1154.8 \text{ J} \]
\[ q_{\text{water}} = -q_{\text{ammonium sulfate}} = 1154.8 \text{ J} = 1.1548 \text{ kJ} \]

$\Delta H$(per mol NH$_4$NO$_3$) = 1.1548kJ/.0484 mol = 23.86 kJ/mol

(b) Endothermic
Bomb Calorimetry

Reactions can be carried out in a sealed “bomb,” such as this one, and measure the heat absorbed by the water.
Bomb Calorimetry

- Because the volume in the bomb calorimeter is constant, what is measured is really the $\Delta E$, not $\Delta H$.

- For most reactions, $\Delta E \approx \Delta H$

- Why?
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.

\[ \Delta H = \Delta E + \Delta PV = \Delta E \]
Hess’s Law

• $\Delta H$ is known for many reactions. Measuring $\Delta H$ can be a pain. Can we estimate $\Delta H$ using $\Delta H$ values for other reactions?
Hess’s Law

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

Yes!

Hess’s law: states that:

ΔH = ΔH₁ + ΔH₂ + ΔH₃

ΔH₁ = -890 kJ
ΔH₂ = -607 kJ
ΔH₃ = -283 kJ
Thermochemistry

Hess’s Law

Why?
Because $\Delta H$ is a state function, and is pathway independent.

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

\[ \text{CH}_4(g) + 2 \text{O}_2(g) \rightarrow \text{CO}(g) + 2 \text{H}_2\text{O}(l) + \frac{1}{2} \text{O}_2(g) \]

\[ \Delta H_1 = -890 \text{ kJ} \]

\[ \Delta H_2 = -607 \text{ kJ} \]

\[ \Delta H_3 = -283 \text{ kJ} \]

\[ \text{CO}_2(g) + 2 \text{H}_2\text{O}(l) \]
Hess’s law, example:

• Given:
  - \( \text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}(\text{g}) \quad \Delta H = 180.7 \text{ kJ} \)
  - \( 2\text{NO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}_2(\text{g}) \quad \Delta H = -113.1 \text{ kJ} \)
  - \( 2\text{N}_2\text{O}(\text{g}) \rightarrow 2\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \Delta H = -163.2 \text{ kJ} \)

• use Hess’s law to calculate \( \Delta H \) for the reaction:
  - \( \text{N}_2\text{O}(\text{g}) + \text{NO}_2(\text{g}) \rightarrow 3\text{NO}(\text{g}) \)
Hess’s law, example:

- Given:
  - \( N_2(g) + O_2(g) \rightarrow 2NO(g) \quad \Delta H = 180.7 \text{ kJ} \)
  - \( 2NO(g) + O_2(g) \rightarrow 2NO_2(g) \quad \Delta H = -113.1 \text{ kJ} \)
  - \( 2N_2O(g) \rightarrow 2N_2(g) + O_2(g) \quad \Delta H = -163.2 \text{ kJ} \)

- Use Hess’s law to calculate \( \Delta H \) for the reaction:
  - \( N_2O(g) + NO_2(g) \rightarrow 3NO(g) \)

\[
\begin{align*}
N_2O(g) & \rightarrow N_2(g) + 1/2O_2(g) \quad \Delta H = -163.2/2 = -81.6 \text{ kJ} \\
NO_2(g) & \rightarrow NO(g) + 1/2O_2(g) \quad \Delta H = 113.1 \text{ kJ}/2 = 56.6 \text{ kJ} \\
N_2(g) + O_2(g) & \rightarrow 2NO(g) \quad \Delta H = 180.7 \text{ kJ} \\
\end{align*}
\]

\( \therefore N_2O(g) + NO_2(g) \rightarrow 3NO(g) \quad \Delta H = 155.7 \text{ kJ} \)
Enthalpies of Formation

An enthalpy of formation, $\Delta H_f$, is defined as the $\Delta H$ for the reaction in which a compound is made from its constituent elements in their elemental forms.
Standard Enthalpies of Formation

Standard enthalpies of formation, $\Delta H_f^\circ$, are measured under standard conditions (25°C and 1.00 atm pressure).

<table>
<thead>
<tr>
<th>Substance</th>
<th>Formula</th>
<th>$\Delta H_f^\circ$ (kJ/mol)</th>
<th>Substance</th>
<th>Formula</th>
<th>$\Delta H_f^\circ$ (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Acetylene</td>
<td>C$_2$H$_2$(g)</td>
<td>226.7</td>
<td>Hydrogen chloride</td>
<td>HCl(g)</td>
<td>−92.30</td>
</tr>
<tr>
<td>Ammonia</td>
<td>NH$_3$(g)</td>
<td>−46.19</td>
<td>Hydrogen fluoride</td>
<td>HF(g)</td>
<td>−268.60</td>
</tr>
<tr>
<td>Benzene</td>
<td>C$_6$H$_6$(l)</td>
<td>49.0</td>
<td>Hydrogen iodide</td>
<td>HI(g)</td>
<td>25.9</td>
</tr>
<tr>
<td>Calcium carbonate</td>
<td>CaCO$_3$(s)</td>
<td>−1207.1</td>
<td>Methane</td>
<td>CH$_4$(g)</td>
<td>−74.80</td>
</tr>
<tr>
<td>Calcium oxide</td>
<td>CaO(s)</td>
<td>−635.5</td>
<td>Methanol</td>
<td>CH$_3$OH(l)</td>
<td>−238.6</td>
</tr>
<tr>
<td>Carbon dioxide</td>
<td>CO$_2$(g)</td>
<td>−393.5</td>
<td>Propane</td>
<td>C$_3$H$_8$(g)</td>
<td>−103.85</td>
</tr>
<tr>
<td>Carbon monoxide</td>
<td>CO(g)</td>
<td>−110.5</td>
<td>Silver chloride</td>
<td>AgCl(s)</td>
<td>−127.0</td>
</tr>
<tr>
<td>Diamond</td>
<td>C(s)</td>
<td>1.88</td>
<td>Sodium bicarbonate</td>
<td>NaHCO$_3$(s)</td>
<td>−947.7</td>
</tr>
<tr>
<td>Ethane</td>
<td>C$_2$H$_6$(g)</td>
<td>−84.68</td>
<td>Sodium carbonate</td>
<td>Na$_2$CO$_3$(s)</td>
<td>−130.9</td>
</tr>
<tr>
<td>Ethanol</td>
<td>C$_2$H$_5$OH(l)</td>
<td>−277.7</td>
<td>Sodium chloride</td>
<td>NaCl(s)</td>
<td>−410.9</td>
</tr>
<tr>
<td>Ethylene</td>
<td>C$_2$H$_4$(g)</td>
<td>52.30</td>
<td>Sucrose</td>
<td>C$<em>12$H$</em>{22}$O$_{11}$(s)</td>
<td>−2221</td>
</tr>
<tr>
<td>Glucose</td>
<td>C$_6$H$_12$O$_6$(s)</td>
<td>−1273</td>
<td>Water</td>
<td>H$_2$O(l)</td>
<td>−285.8</td>
</tr>
<tr>
<td>Hydrogen bromide</td>
<td>HBr(g)</td>
<td>−36.23</td>
<td>Water vapor</td>
<td>H$_2$O(g)</td>
<td>−241.8</td>
</tr>
</tbody>
</table>
Calculation of $\Delta H$

\[ \text{C}_3\text{H}_8 (g) + 5 \text{O}_2 (g) \rightarrow 3 \text{CO}_2 (g) + 4 \text{H}_2\text{O} (l) \]

- Imagine this as occurring in 3 steps:

1. \[ \text{C}_3\text{H}_8 (g) \rightarrow 3 \text{C}_{\text{graphite}} + 4 \text{H}_2 (g) \]
2. \[ \Delta H_1 = +103.85 \text{ kJ} \]
3. \[ \Delta H_2 = -1181 \text{ kJ} \]
4. \[ \Delta H_{\text{rxn}} = -2220 \text{ kJ} \]
5. \[ \Delta H_3 = -1143 \text{ kJ} \]
6. \[ \Delta H_3 = -1143 \text{ kJ} \]
Calculation of $\Delta H$

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

- Imagine this as occurring in 3 steps:

1. $C_3H_8 (g) \rightarrow 3 C\text{(graphite)} + 4 H_2 (g)$
2. $3 C\text{(graphite)} + 3 O_2 (g) \rightarrow 3 CO_2 (g)$
3. $3 CO_2(g) + 4 H_2(g) + 2 O_2(g) \rightarrow 3 CO_2(g) + 4 H_2O(l)$

$\Delta H_1 = +103.85 \text{ kJ}$

$\Delta H_2 = -1181 \text{ kJ}$

$\Delta H_{\text{rxn}} = -2220 \text{ kJ}$

$\Delta H_3 = -1143 \text{ kJ}$
Calculation of $\Delta H$

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

- Imagine this as occurring in 3 steps:

  1. $C_3H_8 (g) \rightarrow 3 C_{\text{(graphite)}} + 4 H_2 (g)$
  2. $3 C_{\text{(graphite)}} + 3 O_2 (g) \rightarrow 3 CO_2 (g)$
  3. $4 H_2 (g) + 2 O_2 (g) \rightarrow 4 H_2O (l)$

$$\Delta H_1 = +103.85 \text{ kJ}$$
$$\Delta H_2 = -1181 \text{ kJ}$$
$$\Delta H_{\text{rxn}} = -2220 \text{ kJ}$$
$$\Delta H_3 = -1143 \text{ kJ}$$
Calculation of $\Delta H$

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

- The sum of these equations is:

$C_3H_8 (g) \rightarrow 3 C_{(\text{graphite})} + 4 H_2 (g)$

$3 C_{(\text{graphite})} + 3 O_2 (g) \rightarrow 3 CO_2 (g)$

$4 H_2 (g) + 2 O_2 (g) \rightarrow 4 H_2O (l)$

$C_3H_8 (g) + 5 O_2 (g) \rightarrow 3 CO_2 (g) + 4 H_2O (l)$
Calculation of $\Delta H$

We can use Hess’s law in this way:

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

where $n$ and $m$ are the stoichiometric coefficients.
Calculation of $\Delta H$

- $C_3H_8 + 5 O_2 \rightarrow 3CO_2 + 4H_2O$

$\Delta H = [3(\Delta H_{f CO_2}) + 4(\Delta H_{f H_2O})] - [(\Delta H_{f C_3H_8}) + (5\Delta H_{f O_2})]$

$= [3(-393.5 \text{ kJ}) + 4(-285.8 \text{ kJ})] - [(-103.85 \text{ kJ}) + 5(0)]$

$= [-1180.5 \text{ kJ} + (-1143.2 \text{ kJ})] - [(-103.85 \text{ kJ}) + 0 \text{ kJ}]$

$= [-2323.7 \text{ kJ}] - [-103.85 \text{ kJ}]$

$= -2219.9 \text{ kJ}$
Energy in Foods

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

<table>
<thead>
<tr>
<th>Approximate Composition (% by mass)</th>
<th>Fuel Value</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>kJ/g  kcal/g (Cal/g)</td>
</tr>
<tr>
<td>Carbohydrate</td>
<td>17  4</td>
</tr>
<tr>
<td>Fat</td>
<td>38  9</td>
</tr>
<tr>
<td>Protein</td>
<td>17  4</td>
</tr>
<tr>
<td>Apples</td>
<td>2.5  0.59</td>
</tr>
<tr>
<td>Beer*</td>
<td>1.8  0.42</td>
</tr>
<tr>
<td>Bread</td>
<td>12  2.8</td>
</tr>
<tr>
<td>Cheese</td>
<td>20  4.7</td>
</tr>
<tr>
<td>Eggs</td>
<td>6.0  1.4</td>
</tr>
<tr>
<td>Fudge</td>
<td>18  4.4</td>
</tr>
<tr>
<td>Green beans</td>
<td>1.5  0.38</td>
</tr>
<tr>
<td>Hamburger</td>
<td>15  3.6</td>
</tr>
<tr>
<td>Milk (whole)</td>
<td>3.0  0.74</td>
</tr>
<tr>
<td>Peanuts</td>
<td>23  5.5</td>
</tr>
</tbody>
</table>

*Beers typically contain 3.5% ethanol, which has fuel value.
What’s the deal with fat?

- $C_nH_{2n}O_n + nO_2 \rightarrow \rightarrow \rightarrow nCO_2 + nH_2O + \text{Energy}$

- $C_nH_{2n}O_2 + n-1H_2O \rightarrow \rightarrow \rightarrow \rightarrow \rightarrow \rightarrow nCO_2 + nH_2O$

  more steps

Fat storage.

It also clogs your arteries.
The vast majority of the energy consumed in this country comes from fossil fuels.