Chapter 5
Thermochemistry

Learning Outcomes:
- Interconvert energy units
- Distinguish between the system and the surroundings in thermodynamics
- Calculate internal energy from heat and work and state sign conventions of these quantities
- Explain the concept of a state function and give examples
- Calculate $\Delta H$ from $\Delta E$ and $P \Delta V$
- Relate $q_p$ to $\Delta H$ and indicate how the signs of $q$ and $\Delta H$ relate to whether a process is exothermic or endothermic
- Use thermochemical equations to relate the amount of heat energy transferred in reactions in reactions at constant pressure ($\Delta H$) to the amount of substance involved in the reaction

- **Energy** is the ability to do work or transfer heat.
- **Thermodynamics** is the study of energy and its transformations.
- **Thermochemistry** is the study of chemical reactions and the energy changes that involve heat.

**Heat**
Energy used to cause the temperature of an object to increase.

**Work**
Energy used to cause an object that has mass to move. $w = F \times d$
Electrostatic potential energy

- The most important form of potential energy in molecules is electrostatic potential energy, $E_{el}$:
  \[ E_{el} = \frac{\kappa Q_1 Q_2}{d} \]
  where $\kappa = 8.99 \times 10^9$ J.m/C$^2$

- Electron charge: $1.602 \times 10^{-19}$ C

- The unit of energy commonly used is the Joule:
  \[ 1 \text{ J} = 1 \frac{\text{kg m}^2}{\text{s}^2} \]

Attraction between ions

- Electrostatic attraction occurs between oppositely charged ions.
- Energy is released when chemical bonds are formed; energy is consumed when chemical bonds are broken.
First Law of Thermodynamics

- Energy can be converted from one form to another, but it is neither created nor destroyed.
- Energy can be transferred between the system and surroundings.
- Chemical energy is converted to heat in grills.
- Sunlight is converted to chemical energy in green plants.
- There are many examples of conversion of energy from one form to another.

System and Surroundings

- The **system** includes the molecules of interest.
- The **surroundings** are everything else.
- In thermochemistry we study the exchange of energy between the system and surroundings.
  - **open system** – matter and energy can be exchanged with the surroundings
  - **closed system** – exchange energy--but not matter--with the surroundings.
  - **isolated system** – neither matter nor energy may be exchanged with 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 $E$.

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

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

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

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

the system released energy to the surroundings.
Thermodynamic Quantities: Three Parts

1) A number, 2) a unit, 3) a sign

– A positive $\Delta E$ results when the system gains energy from the surroundings.
– A negative $\Delta E$ results when the system loses energy to the surroundings.

- When energy is exchanged between the system and the surroundings, it is exchanged as either heat ($q$) or work ($w$).
- That is, $\Delta E = q + w$.

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 released by the system into the surroundings, the process is **exothermic**.

State Functions

The internal energy of a system is independent of the path by which the system achieved that state. **Internal energy, \( E \), is a state function.**
State Functions

$q$ and $w$ are not state functions.

$\Delta E$ is the same whether the battery is shorted out or is discharged by running the fan.

- $q$ and $w$ are different in the two cases.

Work

\[ w = -P \Delta V \]

1 L·atm = 101.3 J

$\text{Zn(s)} + 2 \text{H}^+(aq) \rightarrow \text{Zn}^{2+}(aq) + \text{H}_2(g)$
Example
Calculate the work (in J) associated with the expansion of a gas from 44 mL to 63 mL at a constant pressure of 14 atm.

Enthalpy

- Enthalpy is a thermodynamic function equal to the internal energy plus pressure \( \times \) volume: \( H = E + PV \)

When the system changes at constant pressure, the change in enthalpy, \( \Delta H \), is

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

This can be written

\[
\Delta H = \Delta E + P \Delta V
\]

Since \( \Delta E = q + w \) and \( w = -P \Delta V \), we can substitute these into the enthalpy expression:

\[
\begin{align*}
\Delta H &= \Delta E + P \Delta V \\
\Delta H &= (q + w) - w \\
\Delta H &= q
\end{align*}
\]

The enthalpy change, \( \Delta H \), is defined as the heat gained or lost by the system under constant pressure.

\[\Delta H = q_p\]
Properties of Enthalpy

1. Enthalpy is a state function.
2. Enthalpy is an extensive property.
3. Enthalpy is reversible. The enthalpy change for a reaction is equal in magnitude, but opposite in sign, to $\Delta H$ for the reverse reaction.
4. $\Delta H$ for a reaction depends on the state of the products and the state of the reactants.
Endothermic and Exothermic

\[ \Delta H = H_{\text{final}} - H_{\text{initial}} \text{ or } \Delta H = H_{\text{products}} - H_{\text{reactants}} \]

- A process is **endothermic** when \( \Delta H \) is positive (>0).

- A process is **exothermic** when \( \Delta H \) is negative (<0).

Enthalpies of Reaction

This quantity, \( \Delta H \), is called the **enthalpy of reaction**, or the **heat of reaction**.

A ** thermochemical equation** is an equation for which \( \Delta H \) is given:

\[
\begin{align*}
2 \text{H}_2(g) + \text{O}_2(g) & \rightarrow 2 \text{H}_2\text{O}(l) \quad \Delta H = -483.6 \text{ kJ} \\
\text{H}_2(g) + \frac{1}{2} \text{O}_2(g) & \rightarrow \text{H}_2\text{O}(l) \quad \Delta H = -241.8 \text{ kJ}
\end{align*}
\]

The enthalpy changes assume the coefficients are moles of the substances.
Calorimetry

- **Calorimetry**, the measurement of heat released or absorbed by a chemical reaction.
- A **calorimeter** is the device used to measure heat.
- The quantity of heat transferred by the reaction causes a change in temperature of the solution.

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** \((C\text{ in units of } J/K)\).

\[
C = \frac{q}{\Delta T}
\]

- We define **specific heat capacity** (or simply specific heat; \(C_s\) or \(s\) in units of J/g·K) as the amount of energy required to raise the temperature of 1 g of a substance by 1 K.

- If the amount is one mole, it is the **molar heat capacity**.

\[
C_s = s = \frac{q}{m \times \Delta T}
\]

| TABLE 5.2 Specific Heats of Some Substances at 298 K |
|-----------------|-----------------|-----------------|-----------------|
| **Elements**    | **Specific Heat (J/g-K)** | **Substances** | **Specific Heat (J/g-K)** |
| N\(_2\)(g)      | 1.04             | H\(_2\)O(l)     | 4.18             |
| Al(s)           | 0.90             | CH\(_4\)(g)     | 2.20             |
| Fe(s)           | 0.45             | CO\(_2\)(g)     | 0.84             |
| Hg(l)           | 0.14             | CaCO\(_3\)(s)   | 0.82             |

1.000 g H\(_2\)O
\(T_{	ext{initial}} = 15.5°C\)

4.18 MJ (1 cal)/g of heat

1.000 g H\(_2\)O
\(T_{	ext{final}} = 16.5°C\)
Because the specific heat for water is well known (4.184 J/g·K), we can measure \( q \) for the reaction with this equation:

\[
q_{\text{soln}} = C_s \times m \times \Delta T = -q_{\text{rxn}}
\]

The calorimeter and its contents are the surroundings, so \( q_{\text{soln}} \) is found from the mass, heat capacity, and temperature change.

**Example**

A metal pellet with mass 100.0 g, originally at 88.4 °C, is dropped into 125 g of water originally at 25.1 °C. The final temperature of both the pellet and the water is 31.3 °C.

Calculate the heat capacity \( C \) (in J/°C) and specific heat capacity \( C_s \) (in J/g·°C) of the pellet. The specific heat of water is 4.184 J/g·°C.
**Example**

When 200. g of a AgNO\(_3\) solution mixes with 150. g of NaI solution, 2.93 g of AgI precipitates, and the temperature of the solution rises by 1.34\(^\circ\)C. Assume 350. g of solution and a specific heat capacity of 4.184 J/g\(\cdot\)\(^\circ\)C. Calculate \(\Delta H\) for the following:

\[ \text{Ag}^+(aq) + I^-(aq) \rightarrow \text{AgI(s)} \]

**Bomb Calorimetry**

- Because the volume in the bomb calorimeter is constant, what is measured is really the change in internal energy, \(\Delta E\), not \(\Delta H\).
- For most reactions, the difference is small.
- The heat absorbed (or released) by the water is a very good approximation of the enthalpy change for the reaction.
- \(q_{\text{rxn}} = -C_{\text{cal}} \times \Delta T\)
Example

Hess’s Law

- $\Delta H$ is known for many reactions, but it is inconvenient to measure $\Delta H$ for every reaction in which we are interested.
- However, we can calculate $\Delta H$ using published $\Delta H$ values and the properties of enthalpy.
- Hess’s law states that “If a reaction is carried out in a series of steps, $\Delta H$ for the overall reaction will be equal to the sum of the enthalpy changes for the individual steps.” $\Delta H$ is a state function.
Hess’s Law
Most $\Delta H$ values are labeled $\Delta H^\circ$, and measured under standard conditions
- $P = 1$ atm (but for gases $P = 1$ bar)
- $T =$ usually 298.15 K (25.0 °C)
- Concentration = 1 mol/L

Using Hess’s law - when two or more thermochemical equations are added, the
enthalpy change of the resulting equation is the sum of those for the added equations.

\[
\begin{align*}
C(s) + O_2(g) & \rightarrow CO_2(g) \quad \Delta H = -393.5 \text{ kJ} \\
CO_2(g) & \rightarrow CO(g) + \frac{1}{2}O_2(g) \quad \Delta H = +283.0 \text{ kJ} \\
C(s) + \frac{1}{2}O_2(g) & \rightarrow CO(g) \quad \Delta H = -110.5 \text{ kJ}
\end{align*}
\]

Example
Given the thermochemical equations
\[
\begin{align*}
2WO_2(s) + O_2(g) & \rightarrow 2WO_3(s) \quad \Delta H = -506 \text{ kJ} \\
2W(s) + 3O_2(g) & \rightarrow 2WO_3(s) \quad \Delta H = -1686 \text{ kJ}
\end{align*}
\]
calculate the enthalpy change for:
\[
2W(s) + 2O_2(g) \rightarrow 2WO_2(s)
\]
Enthalpies of Formation

• An enthalpy of formation, \( \Delta H_f \), is defined as the enthalpy change for the reaction in which a compound is made from its constituent elements in their elemental forms.

Standard Enthalpy of Formation

• Only one enthalpy value is needed for each substance, called the standard enthalpy of formation.

• The standard enthalpy of formation is the enthalpy change when one mole of a substance in its standard state is formed from the most stable form of the elements in their standard states.

Standard State

• Enthalpy changes depend on the temperature and pressure at which they are measured
  – When applying Hess’s law, all values must refer to the same conditions of pressure and temperature

• The standard state of a substance at a specified temperature is the pure form at 1 atm pressure
  – Tabulated values for enthalpy refer to the standard state, usually at a temperature of 25°C

<table>
<thead>
<tr>
<th>Substance</th>
<th>Formula</th>
<th>( \Delta H_f ) (kJ/mol)</th>
<th>Substance</th>
<th>Formula</th>
<th>( \Delta H_f ) (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Acetylene</td>
<td>C₂H₄(g)</td>
<td>220.7</td>
<td>Hydrogen chloride</td>
<td>HCl(g)</td>
<td>-92.30</td>
</tr>
<tr>
<td>Ammonia</td>
<td>NH₃(g)</td>
<td>-46.19</td>
<td>Hydrogen fluoride</td>
<td>HF(g)</td>
<td>-268.60</td>
</tr>
<tr>
<td>Benzene</td>
<td>C₆H₆(l)</td>
<td>49.0</td>
<td>Hydrogen iodide</td>
<td>HI(g)</td>
<td>225.9</td>
</tr>
<tr>
<td>Calcium carbonate</td>
<td>CaCO₃(s)</td>
<td>-1287.1</td>
<td>Methane</td>
<td>CH₄(g)</td>
<td>-74.88</td>
</tr>
<tr>
<td>Calcium oxide</td>
<td>CaO(s)</td>
<td>-631.5</td>
<td>Methanol</td>
<td>C₂H₅OH(l)</td>
<td>-238.6</td>
</tr>
<tr>
<td>Carbon dioxide</td>
<td>CO₂(g)</td>
<td>-393.5</td>
<td>Propene</td>
<td>C₃H₆(g)</td>
<td>-109.85</td>
</tr>
<tr>
<td>Carbon monoxide</td>
<td>CO(g)</td>
<td>-110.5</td>
<td>Silver chloride</td>
<td>AgCl(s)</td>
<td>-127.6</td>
</tr>
<tr>
<td>Diamond</td>
<td>C(s)</td>
<td>1.88</td>
<td>Sodium bicarbonate</td>
<td>NaHCO₃(s)</td>
<td>-947.7</td>
</tr>
<tr>
<td>Ethane</td>
<td>C₂H₆(g)</td>
<td>-84.68</td>
<td>Sodium carbonate</td>
<td>Na₂CO₃(s)</td>
<td>-1130.9</td>
</tr>
<tr>
<td>Ethanol</td>
<td>C₂H₅OH(l)</td>
<td>-277.7</td>
<td>Sodium chloride</td>
<td>NaCl(s)</td>
<td>-410.9</td>
</tr>
<tr>
<td>Ethylene</td>
<td>C₂H₄(g)</td>
<td>52.30</td>
<td>Sucrose</td>
<td>C₆H₁₂O₆(s)</td>
<td>-2221</td>
</tr>
<tr>
<td>Glucose</td>
<td>C₆H₁₂O₆(s)</td>
<td>-1273</td>
<td>Water</td>
<td>H₂O(l)</td>
<td>-286.8</td>
</tr>
<tr>
<td>Hydrogen bromide</td>
<td>HBr(g)</td>
<td>-16.23</td>
<td>Water vapor</td>
<td>H₂O(g)</td>
<td>-241.8</td>
</tr>
</tbody>
</table>
Standard Enthalpy of Formation

• The symbol used for standard enthalpy of formation is \( \Delta H^o \), where the \( ^o \) designates standard state.
• The product is always one mole of a single substance.
• The standard enthalpy of formation of the elements in their most stable form is zero.
• Some examples of standard enthalpies of formation:

\[
\begin{align*}
\text{C(graphite) + O}_2(g) & \rightarrow \text{CO}_2(g) \quad \Delta H^o[\text{CO}_2(g)] \\
\text{H}_2(g) + \frac{1}{2}\text{O}_2(g) & \rightarrow \text{H}_2\text{O}(l) \quad \Delta H^o[\text{H}_2\text{O}(l)] \\
2\text{Na}(s) + \text{Se}(s) + 2\text{O}_2(g) & \rightarrow \text{Na}_2\text{SeO}_4(s) \quad \Delta H^o[\text{Na}_2\text{SeO}_4(s)] \\
\text{H}_2(g) & \rightarrow \text{H}_2(g) \quad \Delta H^o[\text{H}_2(g)] = 0
\end{align*}
\]

Enthalpies of Reaction

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

• Imagine this as occurring in 3 steps:

\[
\begin{align*}
\text{C}_3\text{H}_8(g) & \rightarrow 3 \text{C(graphite)} + 4 \text{H}_2(g) \\
3 \text{C(graphite)} + 3 \text{O}_2(g) & \rightarrow 3 \text{CO}_2(g) \\
3 \text{CO}_2(g) + 4 \text{H}_2(g) & \rightarrow 4 \text{H}_2\text{O}(l)
\end{align*}
\]

\[\Delta H = \Sigma n \Delta H^o(\text{products}) - \Sigma m \Delta H^o(\text{reactants})\]

where \( n \) and \( m \) are the stoichiometric coefficients.
C₃H₈ (g) + 5 O₂ (g) → 3 CO₂ (g) + 4 H₂O (l)

Imagine this as occurring in 3 steps:
C₃H₈ (g) → 3 C(graphite) + 4 H₂ (g)
3 C(graphite) + 3 O₂ (g) → 3 CO₂ (g)
4 H₂ (g) + 2 O₂ (g) → 4 H₂O (l)

\[ \Delta H = \sum n \Delta H_f^\circ (products) - \sum m \Delta H_f^\circ (reactants) \]
where \( n \) and \( m \) are the stoichiometric coefficients.

\[ \Delta H = \sum n \Delta H_f^\circ (products) - \sum m \Delta H_f^\circ (reactants) \]
\[ C_3H_8(g) + 5 O_2(g) \rightarrow 3 CO_2(g) + 4 H_2O(l) \]

- The sum of these equations is:

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

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

\[ \Delta H = \sum n \Delta H_f^\circ (\text{products}) - \sum m \Delta H_f^\circ (\text{reactants}) \]

where \( n \) and \( m \) are the stoichiometric coefficients.

\[ \Delta H_{\text{rxn}}^\circ = [(3 \text{ mol})(-393.5 \text{ kJ/mol}) + (4 \text{ mol})(-285.8 \text{ kJ/mol})] - [(1 \text{ mol})(-103.85 \text{ kJ/mol}) + (5 \text{ mol})(0 \text{ kJ})] \]
\[ = [(-1180.5 \text{ kJ}) + (-1143.2 \text{ kJ})] - [(-103.85 \text{ kJ}) + (0 \text{ kJ})] \]
\[ = (-2323.7 \text{ kJ}) - (-103.85 \text{ kJ}) \]
\[ = -2219.9 \text{ kJ} \]
**Example**

Use standard enthalpies of formation to calculate the enthalpy change for the reaction:

\[ \text{P}_4\text{O}_{10}(s) + 6\text{H}_2\text{O}(g) \rightarrow 4\text{H}_3\text{PO}_4(s) \]

<table>
<thead>
<tr>
<th>Substance</th>
<th>( \Delta H_f^\circ ) (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{P}<em>4\text{O}</em>{10}(s) )</td>
<td>-2940</td>
</tr>
<tr>
<td>( \text{H}_2\text{O}(g) )</td>
<td>-242</td>
</tr>
<tr>
<td>( \text{H}_3\text{PO}_4(s) )</td>
<td>-1279</td>
</tr>
</tbody>
</table>

\[ \Delta H = \Sigma n \Delta H_f^\circ \text{ (products)} - \Sigma m \Delta H_f^\circ \text{ (reactants)} \]

where \( n \) and \( m \) are the stoichiometric coefficients.

**Bond Enthalpy**

The enthalpy associated with breaking one mole of a particular bond in a gaseous substance.

\[ \text{Cl}_2(g) \rightarrow 2\text{Cl}(g) \]

\[ \text{CH}_4(g) \rightarrow \text{C}(g) + 4\text{H}(g) \]
• Bond enthalpy is always positive because energy is required to break chemical bonds.
• Energy is released when a bond forms between gaseous fragments.
• The greater the bond enthalpy, the stronger the bond.

### Table 5.4 Average Bond Enthalpies (kJ/mol)

<table>
<thead>
<tr>
<th>Bond</th>
<th>Enthalpy (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>C—H</td>
<td>413</td>
</tr>
<tr>
<td>N—H</td>
<td>391</td>
</tr>
<tr>
<td>O—H</td>
<td>463</td>
</tr>
<tr>
<td>F—F</td>
<td>155</td>
</tr>
<tr>
<td>C—C</td>
<td>348</td>
</tr>
<tr>
<td>N—N</td>
<td>163</td>
</tr>
<tr>
<td>O—O</td>
<td>495</td>
</tr>
<tr>
<td>Cl—F</td>
<td>253</td>
</tr>
<tr>
<td>C—N</td>
<td>293</td>
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<tr>
<td>N—F</td>
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<tr>
<td>O—F</td>
<td>190</td>
</tr>
<tr>
<td>Cl—Cl</td>
<td>242</td>
</tr>
<tr>
<td>C—O</td>
<td>358</td>
</tr>
<tr>
<td>N—Cl</td>
<td>200</td>
</tr>
<tr>
<td>O—Cl</td>
<td>203</td>
</tr>
<tr>
<td>C—O</td>
<td>799</td>
</tr>
<tr>
<td>N—Br</td>
<td>243</td>
</tr>
<tr>
<td>O—I</td>
<td>234</td>
</tr>
<tr>
<td>Br—I</td>
<td>237</td>
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<tr>
<td>C—F</td>
<td>485</td>
</tr>
<tr>
<td>Br—Cl</td>
<td>218</td>
</tr>
<tr>
<td>C—Cl</td>
<td>328</td>
</tr>
<tr>
<td>H—H</td>
<td>436</td>
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<tr>
<td>Br—Br</td>
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<tr>
<td>C—Br</td>
<td>276</td>
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<tr>
<td>H—F</td>
<td>567</td>
</tr>
<tr>
<td>C—I</td>
<td>240</td>
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<tr>
<td>H—Cl</td>
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<tr>
<td>I—Cl</td>
<td>208</td>
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<td>C—I</td>
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<td>H—Br</td>
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<tr>
<td>C—I</td>
<td>240</td>
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<tr>
<td>H—I</td>
<td>299</td>
</tr>
<tr>
<td>I—I</td>
<td>151</td>
</tr>
</tbody>
</table>

### Bond Enthalpies and Enthalpy of Reaction

To obtain an estimate of $\Delta H$, 
• Sum the bond enthalpies for all bonds broken and subtract the sum of the bond enthalpies for all bonds formed.

$$\Delta H_{\text{rxn}} = \Sigma (\text{bond enthalpies of bonds broken}) - \Sigma (\text{bond enthalpies of bonds formed})$$
Predict whether a chemical reaction will be endothermic or exothermic using bond enthalpies.

**Energy in Foods**

The energy released when one gram of food is combusted is its **fuel value**.

<table>
<thead>
<tr>
<th>TABLE 5.5 Compositions and Fuel Values of Some Common Foods</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Approximate Composition (% by Mass)</strong></td>
</tr>
<tr>
<td></td>
</tr>
<tr>
<td>Carbohydrate</td>
</tr>
<tr>
<td>Fat</td>
</tr>
<tr>
<td>Protein</td>
</tr>
<tr>
<td>Apples</td>
</tr>
<tr>
<td>Beer*</td>
</tr>
<tr>
<td>Bread</td>
</tr>
<tr>
<td>Cheese</td>
</tr>
<tr>
<td>Eggs</td>
</tr>
<tr>
<td>Fudge</td>
</tr>
<tr>
<td>Green beans</td>
</tr>
<tr>
<td>Hamburger</td>
</tr>
<tr>
<td>Milk (whole)</td>
</tr>
<tr>
<td>Peanuts</td>
</tr>
</tbody>
</table>

* Beer typically contains 3.5% ethanol, which has fuel value.
• Most of the energy in foods comes from carbohydrates, fats, and proteins.

• Carbohydrates (17 kJ/g):
  \[ C_6H_{12}O_6(s) + 6 O_2(g) \rightarrow 6 CO_2(g) + 6 H_2O(l) \quad \Delta H^\circ = -2803 \text{ kJ} \]

• Fats (38 kJ/g):
  \[ 2 C_{37}H_{110}O_6(s) + 163 O_2(g) \rightarrow 114 CO_2(g) + 110 H_2O(l) \quad \Delta H^\circ = -71,609 \text{ kJ} \]

• Proteins produce 17 kJ/g (same as carbohydrates):
  However, their chemical reaction in the body is NOT the same as in a calorimeter.

---

**Fuels**

<table>
<thead>
<tr>
<th>TABLE 5.6 Fuel Values and Compositions of Some Common Fuels</th>
</tr>
</thead>
<tbody>
<tr>
<td>Approximate Elemental Composition (Mass %)</td>
</tr>
<tr>
<td>-------------------------------------------</td>
</tr>
<tr>
<td>Wood (pine)</td>
</tr>
<tr>
<td>Anthracite coal (Pennsylvania)</td>
</tr>
<tr>
<td>Bituminous coal (Pennsylvania)</td>
</tr>
<tr>
<td>Charcoal</td>
</tr>
<tr>
<td>Crude oil (Texas)</td>
</tr>
<tr>
<td>Gasoline</td>
</tr>
<tr>
<td>Natural gas</td>
</tr>
<tr>
<td>Hydrogen</td>
</tr>
</tbody>
</table>

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