**KE and PE**

- **Potential energy** is energy of position
- **Kinetic Energy** is energy of motion
  - the energy of a moving object
    \[ KE = \frac{1}{2} \text{mass} \times \text{velocity}^2 \]
- Potential energy can be converted into kinetic energy.
- **Missing Energy** is:

**Attractive Forces**

Energy is released when two objects with an attractive force are allowed to come together

Fundamental attractive forces:

1. 
2. 
3. 

**Units of energy**

SI Unit for energy is the Joule, J:

\[ KE = \frac{1}{2} mv^2 \]
\[ = \frac{1}{2} (2 \text{ kg})^2 (1 \text{ m/s})^2 \]
\[ = 1 \text{ kg m}^2/\text{s}^2 \]
\[ = 1 \text{ J} \]

Old unit: calorie (before Joule)

amount of heat req. to warm 1.0 g water 1 degree C

1 cal = ______________________ (exactly)

A nutritional Calorie:

1 Cal = ______________________

**Systems and Surroundings**

- System: ___________________________
- Surroundings: ___________________________

**Work, Energy Heat and Temperature**

- Force is a push or pull on an object.
- **Work** is force applied to object over a distance:
  - \[ W = F \times d \]
- Work is related to pressure \[ W = -\Delta PV \]
Energy

- Energy is ability to make things change
  - Energy does work (mechanical E)
  - Total internal energy (KE, PE, others)
  - Heat (thermal E, transferred between objects)
  - Heat energy change = $q$

- Temperature is related to the _____________ of an object

First Law of Thermo

- Energy cannot be created or destroyed.
- Energy of (system + surroundings) is constant.

$$\Delta E = \Delta E_{\text{sys}} + \Delta E_{\text{sur}}$$

Energy and chemistry

- 2 things energy changes in chemical reactions can do:
  1. change __________________ (hotter or colder)
  2. ____________ (P V)

- Energy goes from system to surroundings (and vice versa).

$$\Delta E_{\text{sys}} + \Delta E_{\text{sur}} = \text{____________}$$

$$\Delta E_{\text{sys}} = -\Delta E_{\text{sur}}$$

Absolute Energy of a system can’t be measured, must measure __________________

State Functions

- **State function:** depends only on diff of initial and final states of system, not on the path to get there
- **Path function:** depends on path taken
- Change in altitude up a mountain is a ____________
- Distance that one travels going up a mountain is a ____________
- The balance in your bank account is a ____________

Concept Check:
Is $\Delta E$ a state or path function?
• Endothermic: reactants ________________ heat from the surroundings. feels ________________
• Exothermic: reactants ________________ heat to the surroundings. feels ________________

At constant pressure, the heat energy of a substance is called ________________ = H

ΔH is change in enthalpy of system before and after a chemical reaction:

ΔH = H_{final} - H_{initial}

Is enthalpy a state function? Or a path function?

ΔH = + or -? Endo or exothermic?

Enthalpy

ΔH = _______: system (chemicals) gains heat from surroundings

ΔH = _______: system loses heat to surroundings

Why do some reactions give off heat?

Why exothermic?

For a system at constant pressure,

\[ \Delta H = q \]

So, at constant pressure, the change in enthalpy is the heat gained or lost.

Heat transfer = q = ____________________________
Specific Heat Capacity (s)  Section 5.5

A property of matter that determines the Temp change when heat is applied to a substance.

Listed in J/g °C or J/g K (same numerical value).

Describes how much E req. to raise 1 g by 1 °C.

Which takes more energy to raise temp by 1 °C?

a. 1.0 g Fe  s = 0.44 J/g °C
b. 1.0 g Al  s = 0.90 J/g °C

Which gets hotter with 100 J energy applied?

A 125 g piece of Cu metal (s = 0.387 J/g °C) is heated from 25.0 to 312 °C. How much heat does the Cu absorb?

How much water can be heated from 25.0 to 50.0 °C by the amount of E calculated above?

How do chemists measure ∆H?

(We can calculate from standard tables too)

1. Constant Pressure  (q = ∆H)
2. Heat transfer = mass x ∆T x specific heat
3. q = m x ∆T x s

In calorimeter, a chemical reaction may release energy. E released by reactants is absorbed by water. What must be the sum of E released + E absorbed?

Due to 1st law of thermo:
Heat absorbed by H₂O + heat released by rxn =

Or:  q_water = -q_reaction

50.0 mL of 0.500 M NaOH is added to 25.0 mL of 0.500 M HCl in a calorimeter. Both are at an initial temp. of 25.00 °C, and the final temp is 27.21 °C.

Assume: the mass of the final solution is 76.0 g and the s of the solution is 4.18 J/g °C

What is the q of water?
What is the q of the reaction?

What is the ∆H of reaction, in kJ/mol?
**Bomb Calorimetry (Constant Volume)**

- Reaction done under constant volume.
- Study combustion.

**Thermochemical equations**

Section 5.4

1. A chemical equation plus the value of $\Delta H$
2. Can be treated like math equations (add, subtract)
3. Fractions are OK
4. $\Delta H$ is proportional to amount of reactants:

   H$_2$ + $\frac{1}{2}$ O$_2$ $\rightarrow$ H$_2$O

5. If equation is reversed, change the sign of $\Delta H$:

   CO$_2$(g) + 2H$_2$O(g) $\rightarrow$ CH$_4$(g) + 2O$_2$(g) $\Delta H = +802$ kJ

6. States of matter matter!

   H$_2$O(g) $\rightarrow$ H$_2$O(l) $\Delta H = -44$ kJ

**Math review:**

What can you do when adding the following eq?

$$x + y = z$$
$$x + z = A$$

Add together, cancel on both sides

$$2x + y = A$$

The same can be done with thermochem eq.

**Hess’s Law**

- Used to calculate $\Delta H$ instead of doing exp.
- $\Delta H$ for overall reaction = sum of $\Delta H$ for each step.

Example:

$$\text{CH}_4(g) + 2\text{O}_2(g) \rightarrow \text{CO}_2(g) + 2\text{H}_2\text{O}(g) \quad \Delta H = -802 \text{ kJ}$$

Let condense

$$2\text{H}_2\text{O}(g) \rightarrow 2\text{H}_2\text{O}(l) \quad \Delta H = -88 \text{ kJ}$$

Add together, cancel on both sides

$$\text{CH}_4(g) + 2\text{O}_2(g) \rightarrow \text{CO}_2(g) + 2\text{H}_2\text{O}(l) \quad \Delta H = -890 \text{ kJ}$$

$$\text{CH}_4(g) + 2\text{O}_2(g) \rightarrow \text{CO}_2(g) + 2\text{H}_2\text{O}(g) \quad \Delta H = -820 \text{ kJ}$$

CO$_2$(g) + $\frac{1}{2}$ O$_2$(g) $\rightarrow$ CO(g) $\Delta H = -283$ kJ

CH$_4$(g) + 3/2O$_2$(g) $\rightarrow$ CO(g) + 2H$_2$O(l) $\Delta H = \ ?$ kJ
Note that: \[ \Delta H_1 = \Delta H_2 + \Delta H_3 \]

\[ \Delta H_1 = -880 \text{ kJ} \]

\[ \Delta H_2 = -283 \text{ kJ} \]

\[ \Delta H_3 = \frac{1}{2} \Delta H_2 \]

**Enthalpies of Formation**

- enthalpy of formation equation, \( \Delta H_f \): =

  Heat change for 1 mol of a compound formed from its elements, standard conditions (1 atm and 25°C).

- For \( \text{CO}_2(g) \), \( \Delta H_f \) =

  \[ \Delta H_f = -393.5 \text{ kJ/mol} \]

  For \( \text{C}_6\text{H}_7\text{N}(l) \), \( \Delta H_f \) =

**TABLE 5.1 Standard Enthalpies of Formation, \( \Delta H_f \), at 298 K**

<table>
<thead>
<tr>
<th>Substance</th>
<th>Formula</th>
<th>( \Delta H_f ) (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Acetylene</td>
<td>( \text{C}_2\text{H}_2\text{O}(g) )</td>
<td>226.7</td>
</tr>
<tr>
<td>Ammonia</td>
<td>( \text{NH}_3(g) )</td>
<td>-46.19</td>
</tr>
<tr>
<td>Benzene</td>
<td>( \text{C}_2\text{H}_6(l) )</td>
<td>49.0</td>
</tr>
<tr>
<td>Calcium carbonate</td>
<td>( \text{CaCO}_3\text{(s)} )</td>
<td>-1207.1</td>
</tr>
<tr>
<td>Calcium oxide</td>
<td>( \text{CaO}(s) )</td>
<td>-635.5</td>
</tr>
<tr>
<td>Carbon dioxide</td>
<td>( \text{CO}_2(g) )</td>
<td>-393.5</td>
</tr>
<tr>
<td>Carbon monoxide</td>
<td>( \text{CO}(g) )</td>
<td>-110.5</td>
</tr>
<tr>
<td>Diamond</td>
<td>( \text{C}(s) )</td>
<td>1.88</td>
</tr>
<tr>
<td>Ethane</td>
<td>( \text{C}_2\text{H}_2\text{O}(g) )</td>
<td>-84.68</td>
</tr>
<tr>
<td>Ethanol</td>
<td>( \text{C}_2\text{H}_5\text{O}(l) )</td>
<td>-277.7</td>
</tr>
<tr>
<td>Ethylene</td>
<td>( \text{C}_2\text{H}_4(g) )</td>
<td>52.20</td>
</tr>
<tr>
<td>Glucose</td>
<td>( \text{C}_6\text{H}_12\text{O}_6(s) )</td>
<td>-1273</td>
</tr>
<tr>
<td>Hydrogen bromide</td>
<td>( \text{HBr}(g) )</td>
<td>-36.23</td>
</tr>
</tbody>
</table>

**Enthalpies of Formation**

- most stable form of element is used (graphite vs. diamond)

- What is \( \Delta H_f \) equation for \( \text{O}_2 \)?

- Memorize all elemental \( \Delta H_f \)

**Enthalpies of Formation**

- Use \( \Delta H_f \) to calculate enthalpies of a reaction

- Sum \( \Delta H_f \) products – sum \( \Delta H_f \) reactants = \( \Delta H \)

- Works because of Hess’ Law

**Reaction:** \( \text{CO} + \text{N}_2\text{O} \rightarrow \text{CO}_2 + \text{N}_2 \)

\[ \Delta H \]

\[ \Delta H_f \] reactants

\[ \Delta H_f \] products

\[ \Delta H_f \] direct reaction

\[ \Delta H_f \] total reaction

\[ \Delta H_f \] final products
Foods

- 1 nutritional Calorie, 1 Cal = 1000 cal = 1 kcal.
- Energy supply in our bodies comes from carbohydrates and fats (mostly).
- Intestines: carbohydrates converted into glucose:
  \[ C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O, \Delta H = -2816 \text{ kJ} \]
- Fats break down as follows:
  \[ 2C_{57}H_{110}O_6 + 163O_2 \rightarrow 114CO_2 + 110H_2O, \Delta H = -75,520 \text{ kJ} \]

Fuels

- In 2000 the United States consumed \(1.03 \times 10^{17}\) kJ of fuel.
- Most from petroleum and natural gas.
- Remainder from coal, nuclear, and hydroelectric.
- Fossil fuels are not renewable on our time scale.
What we pay for in a gallon of gas  8-06  (DOE)
Assuming $3.00/gal

<table>
<thead>
<tr>
<th>Component</th>
<th>Price Percentage</th>
</tr>
</thead>
<tbody>
<tr>
<td>Federal + state taxes</td>
<td>15%</td>
</tr>
<tr>
<td>Distribution + marketing</td>
<td>14%</td>
</tr>
<tr>
<td>Refining</td>
<td>14%</td>
</tr>
<tr>
<td>Raw materials (crude, additives)</td>
<td>55%</td>
</tr>
</tbody>
</table>

- Hydrogen has a high fuel value: 142 kJ/g.
- What are products of $H_2$ combustion?
- Where does $H_2$ come from?
- Is it a good energy SOURCE?

Hydrogen in $O_2$ in

$H_2$ in $O_2$ in

$e^-$ stripped from $H_2$
Travel through load,
Combine with $O_2$ to make water

Water out

Biodiesel and the Carbon Cycle

Reduced Smog and $CO_2$
Fat + NaOH + methanol

\[ O \quad \text{Fat} \quad \text{NaOH} \quad \text{methanol} \]

\[ \text{CH}_2\text{O} - \text{C} - \text{R}_1 \]
\[ \text{O} \quad \text{R}_1 - \text{C} - \text{O} - \text{CH}_3 \quad \text{CH}_2\text{O} \]

\[ \text{CH} - \text{O} - \text{C} - \text{R}_2 + 3\text{CH}_3\text{OH} \rightarrow \text{R}_2 - \text{C} - \text{O} - \text{CH}_3 \quad \text{CH}_2\text{O} \]

\[ \text{CH}_2\text{O} - \text{C} - \text{R}_3 \]
\[ \text{R}_3 - \text{C} - \text{O} - \text{CH}_3 \quad \text{CH}_2\text{O} \]

75 million gallons of biodiesel 2005

---

### Gasoline Gallon Equivalent (GGE) Table

<table>
<thead>
<tr>
<th>Fuel Type</th>
<th>Unit of Measure</th>
<th>BTUs Per Unit</th>
<th>Gallon Equivalent</th>
</tr>
</thead>
<tbody>
<tr>
<td>Gasoline, regular unleaded, (typical)</td>
<td>gallon</td>
<td>114,100</td>
<td>1.60 gallons</td>
</tr>
<tr>
<td>Gasoline, RFS. (10% B87)</td>
<td>gallon</td>
<td>112,000</td>
<td>1.62 gallons</td>
</tr>
<tr>
<td>Diesel, (typical)</td>
<td>gallon</td>
<td>129,800</td>
<td>0.99 gallons</td>
</tr>
<tr>
<td>Liquid natural gas (LNG), (typical)</td>
<td>gallon</td>
<td>73,000</td>
<td>1.52 gallons</td>
</tr>
<tr>
<td>Compressed natural gas (CNG), (typical)</td>
<td>cubic feet</td>
<td>390</td>
<td>126.67 cu. ft.</td>
</tr>
<tr>
<td>Liquid petroleum gas (LPG or propane)</td>
<td>gallon</td>
<td>84,300</td>
<td>1.28 gallons</td>
</tr>
<tr>
<td>Methanol (M-100)</td>
<td>gallon</td>
<td>25,200</td>
<td>2.01 gallons</td>
</tr>
<tr>
<td>Methanol (M-85)</td>
<td>gallon</td>
<td>65,400</td>
<td>1.74 gallons</td>
</tr>
<tr>
<td>Ethanol (E-100)</td>
<td>gallon</td>
<td>76,100</td>
<td>1.90 gallons</td>
</tr>
<tr>
<td>Ethanol (E-85)</td>
<td>gallon</td>
<td>81,300</td>
<td>1.40 gallons</td>
</tr>
<tr>
<td>Bio-Diesel (B-20)</td>
<td>gallon</td>
<td>125,300</td>
<td>0.80 gallons</td>
</tr>
<tr>
<td>Electricity</td>
<td>kilowatt hour</td>
<td>3,400</td>
<td>33.33 kilowatt</td>
</tr>
</tbody>
</table>