Ch06

Energy

Thermochemistry,
understanding energy, heat & work.

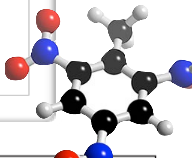
version 1.5

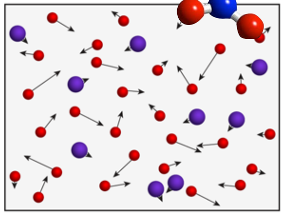
Ch06

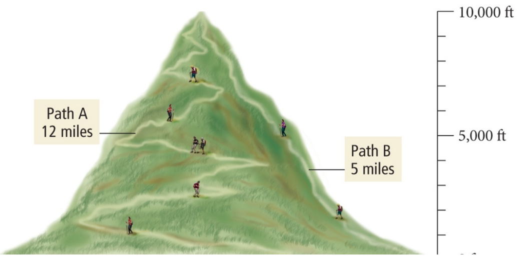
Accounting for Energy

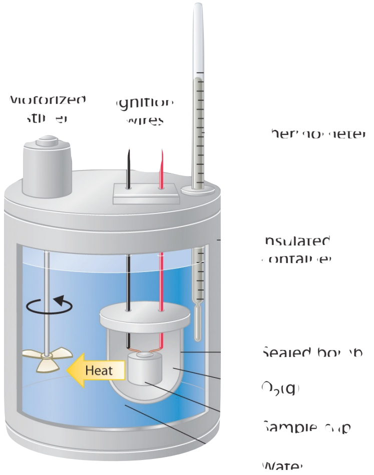
- Energy
  - Definitions
  - Classifications
    - Kinetic, Potential, Thermal & Chemical
  - Units
- Systems – keeping perspective consistent is essential!
  - the system & the surroundings
  - types of systems
    - open, closed, isolated
- Changes, Values & Size
  - What is change? (over time)
  - Measuring changes
    - State Functions vs Path Functions
  - Changes in Energy
    - Internal Energy
    - Heat & Work

- Energy Transfer/Change
  - by Heat
    - Heat (Temperature Change)
    - Heat Capacity
      - Specific Heat Capacity
      - Molar Heat Capacity
    - Heat Transfer
  - by Work
    - Pressure-Volume Work
  - Measuring Changes in Internal Energy
    - Bomb Calorimetry



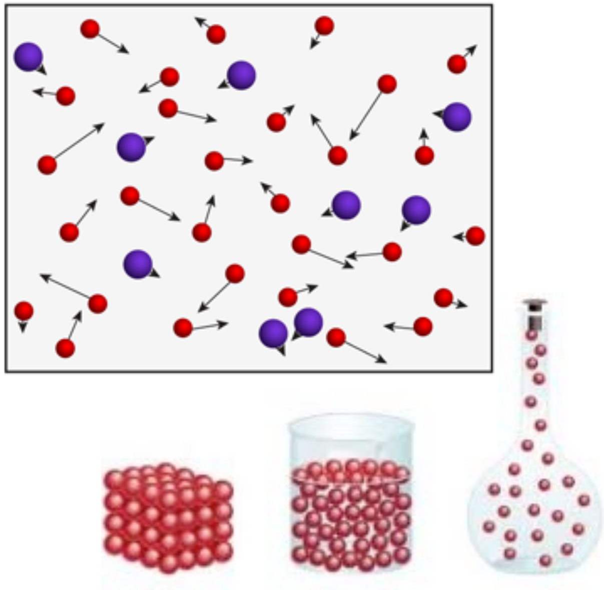






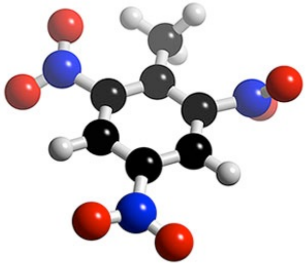
Energy: Kinetic & Potential

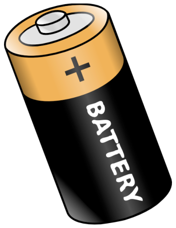
- Energy (E) is the capacity to do work or transmit heat.
  - All energy is kinetic or potential.
  - Kinetic Energy is energy of motion
    - swinging a bat, moving a piston, rolling a ball
    - The movement of particles is thermal energy.
      - Thermal energy is a kind of kinetic energy.
  - Potential Energy is energy of position
    - the position of the edge of a cliff, top of a hill, of a bow string, rubber band, of ions in a battery.
    - The position of atoms in a molecule is chemical energy.
      - Chemical energy is a kind of potential energy.

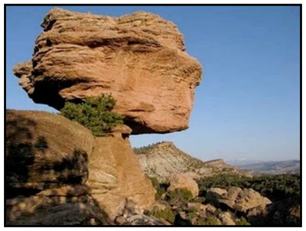


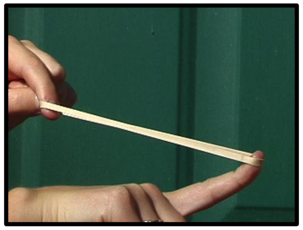
Energy: Kinetic & Potential

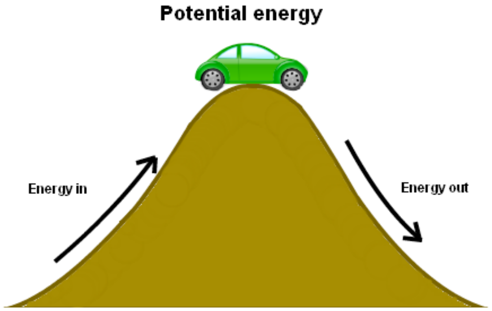
- Energy (E) is the capacity to do work or transmit heat.
  - All energy is kinetic or potential.
  - Kinetic Energy is energy of motion
    - swinging a bat, moving a piston, rolling a ball
    - The movement of particles is thermal energy.
      - Thermal energy is a kind of kinetic energy.
  - Potential Energy is energy of position
    - the position of the edge of a cliff, top of a hill, of a bow string, rubber band, of ions in a battery.
    - The position of atoms in a molecule is chemical energy.
      - Chemical energy is a kind of potential energy.

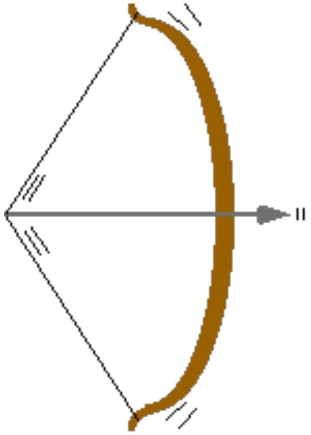


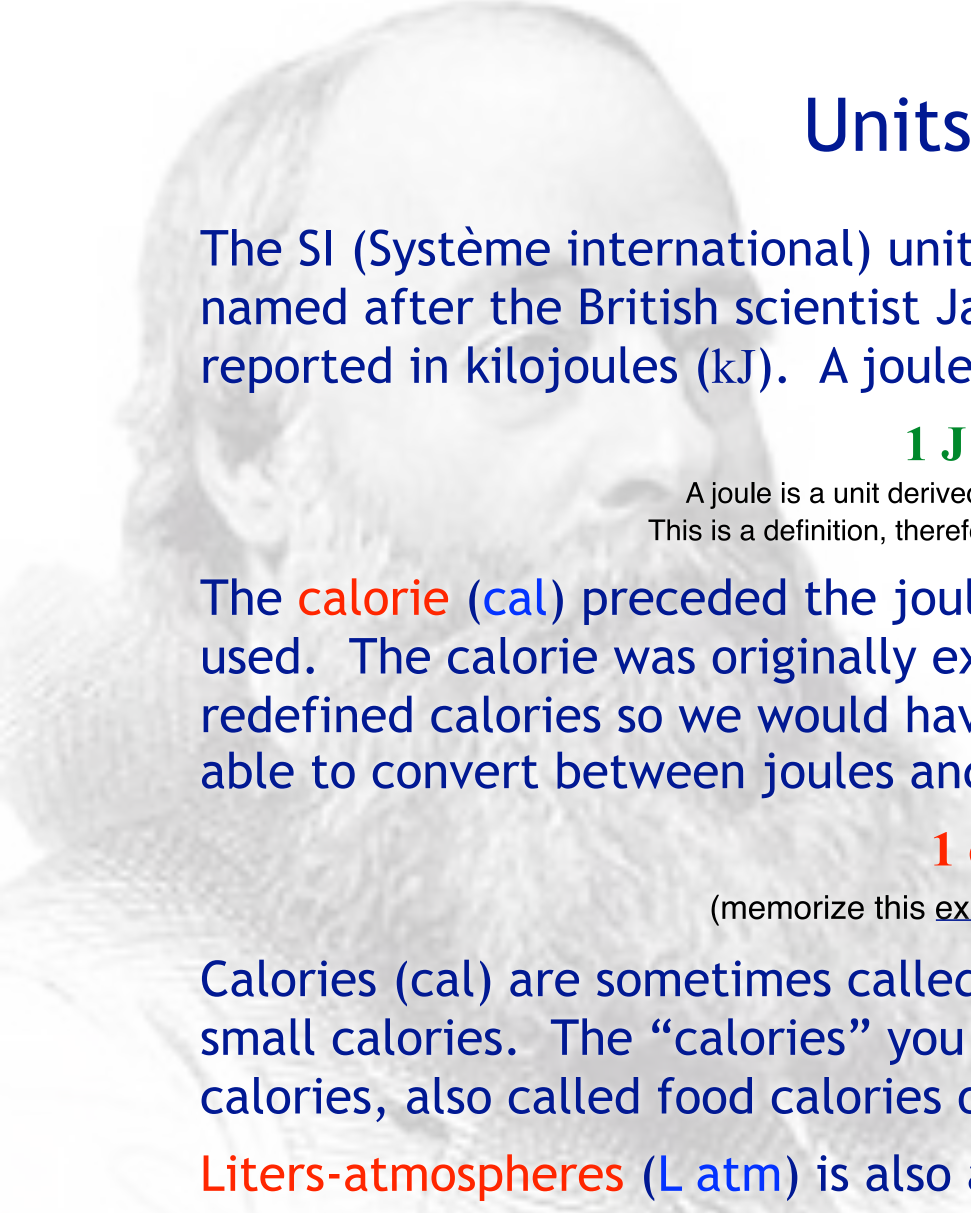












Units of Energy

The SI (Système international) unit of energy is the joule (J), a derived unit named after the British scientist James Joule. It's common to find energy reported in kilojoules (kJ). A joule is defined as:

$$1\ \mathrm{J} = 1\ \mathrm{kg\ m^2 / s^2}$$

A joule is a unit derived from kilograms, meters, and seconds.
This is a definition, therefore an exact conversion factor - ∞ sig figs.

The calorie (cal) preceded the joule as a standard of energy and is still widely used. The calorie was originally experimentally determined. We later redefined calories so we would have an exact conversion factor. You must be able to convert between joules and calories.

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

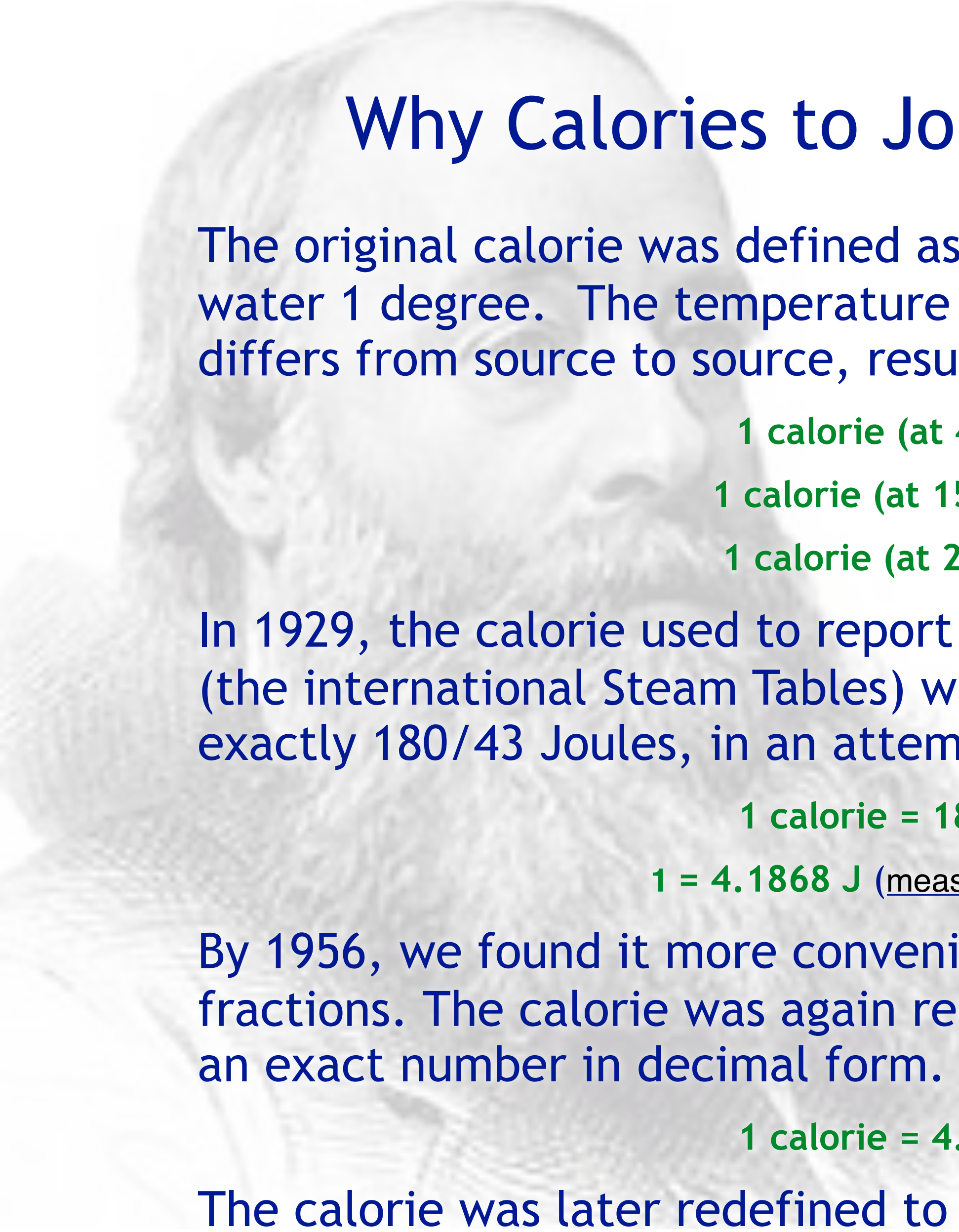
(memorize this exact conversion factor - ∞ sig figs)

Calories (cal) are sometimes called small calories. A large calorie (Cal) is 1000 small calories. The "calories" you see on supermarket packages are large calories, also called food calories or kilocalories.

Liters-atmospheres (L atm) is also a unit of energy. You won't have to do this conversion but I'll use it in some derivations.

$$1\ \mathrm{L\ atm} = 101.325\ \mathrm{J}$$

(you do not need to memorize this measured conversion factor, it will be provided if needed - 6 sig figs)



Why Calories to Joules is an Exact Number

The original calorie was defined as the energy it takes to raise 1 grams of water 1 degree. The temperature at which that experiment is said to occur differs from source to source, resulting in different conversion factors.

1 calorie (at 4°C) = 4.204 J (measured)

1 calorie (at 15°C) = 4.1855 J (measured)

1 calorie (at 20°C) = 4.182 J (measured)

In 1929, the calorie used to report internationally accepted boiling point data (the international Steam Tables) was standardized by defining it to be equal to exactly 180/43 Joules, in an attempt to remove that uncertainty.

1 calorie = 180/43 J (exact - a definition)

1 = 4.1868 J (measured, the fraction rounded to 5 s.f.)

By 1956, we found it more convenient to do calculations with decimals than fractions. The calorie was again redefined to be exactly 4.1868 J to make it an exact number in decimal form.

1 calorie = 4.1868 J (exact - a definition)

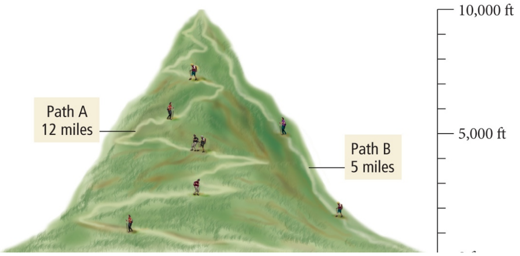
The calorie was later redefined to be equal to exactly 4.184 J and this calorie is now used for all thermochemical tables.

1 calorie = 4.184 J (exact - a definition)

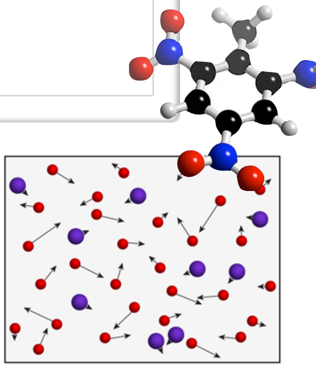
This final calorie is called the Thermochemical calorie,
this is the calorie we will use in this class.

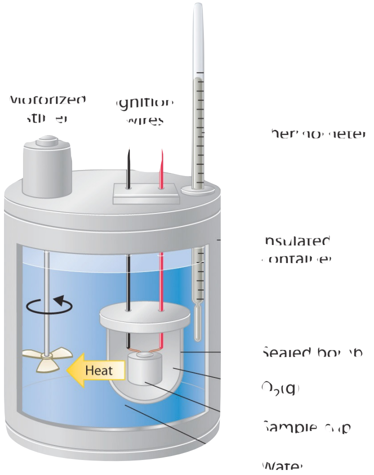
Accounting for Energy

- Energy
  - Definitions
  - Classifications
    - Kinetic, Potential, Thermal & Chemical
  - Units
- Systems – keeping perspective consistent is essential!
  - the system & the surroundings
  - types of systems
    - open, closed, isolated
- Changes, Values & Size
  - What is change? (over time)
  - Measuring changes
    - State Functions vs Path Functions
  - Changes in Energy
    - Internal Energy
    - Heat & Work

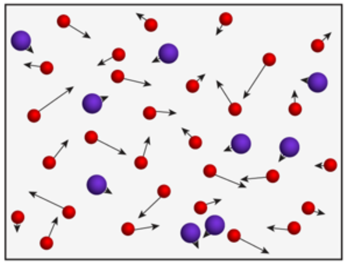


- Energy Transfer/Change
  - by Heat
    - Heat (Temperature Change)
    - Heat Capacity
      - Specific Heat Capacity
      - Molar Heat Capacity
    - Heat Transfer
  - by Work
    - Pressure-Volume Work
  - Measuring Changes in Internal Energy
    - Bomb Calorimetry

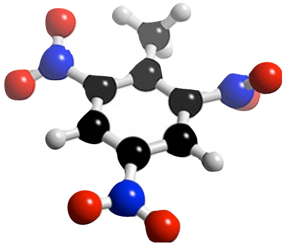




Internal Energy



- A system is any portion of the universe we single out for study.
- Internal Energy is the energy contained within that system.
  - Chemists care about:
    - the thermal energy (kinetic energy)
    - the chemical energy (potential energy)

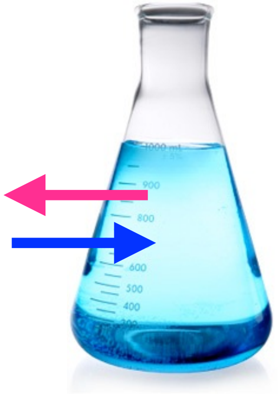


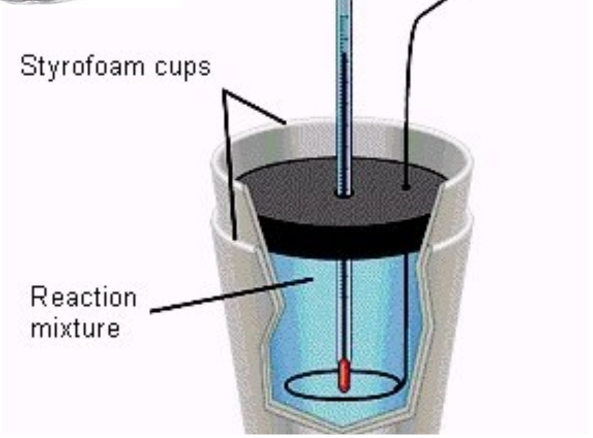
- Not all systems have the same energy.
  - A cold brick may have less energy than a hot brick.
  - A hot brick may have less energy than a block of wood.
  - A block of wood may have less energy than a lump of coal.
  - A lump of coal might have less energy than a vial of nitroglycerine.
- We want to understand how to predict and effect the energy of a system.
  - We want to know what substances may be richer in internal energy.
  - We want to discover how to unlock energy and put it to use.
  - We want to understand how to trap energy in systems to store it for later.

- We want to understand how to measure and control changes to a systems energy.

System and Surroundings

- A system is any portion of the universe we single out for study.
- The surroundings are the rest of the universe (anything not in the system we have selected).
  - Anything that leaves the system, must enter the surroundings.
- A open system is one where matter and energy may both exchange between the system and it's surroundings.
- In a closed system matter cannot exchange, but energy can exchange.
- An isolated system is one where neither may exchange between the system and it's surroundings. Isolated systems are effectively a universe of their own (at least it's our goal to approximate that).
- When we do thermochemical experiments we will do our experiments in isolated apparatus.
- Always understand the boundaries of the system you're looking at.
  - If the system is not isolated, you have to account for all energy gained or lost by that system to it's surroundings — all energy that crosses those boundaries.

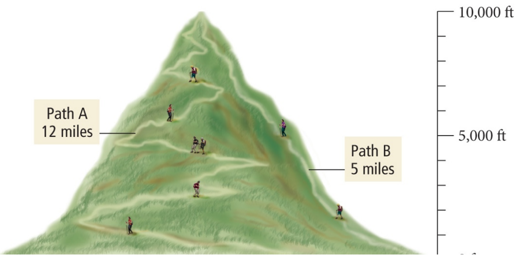




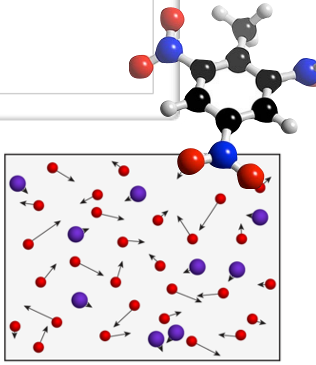
Ch06

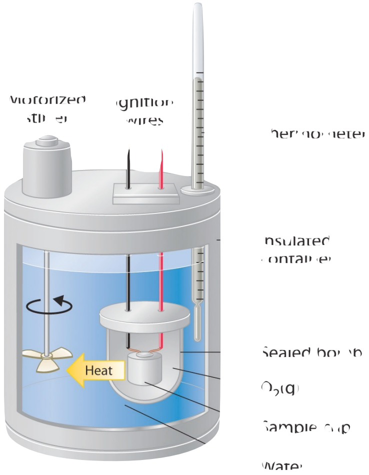
Accounting for Energy

- Energy
  - Definitions
  - Classifications
    - Kinetic, Potential, Thermal & Chemical
  - Units
- Systems — keeping perspective consistent is essential!
  - the system & the surroundings
  - types of systems
    - open, closed, isolated
- Changes, Values & Size
  - What is change? (over time)
  - Measuring changes
    - State Functions vs Path Functions
  - Changes in Energy
    - Internal Energy
    - Heat & Work



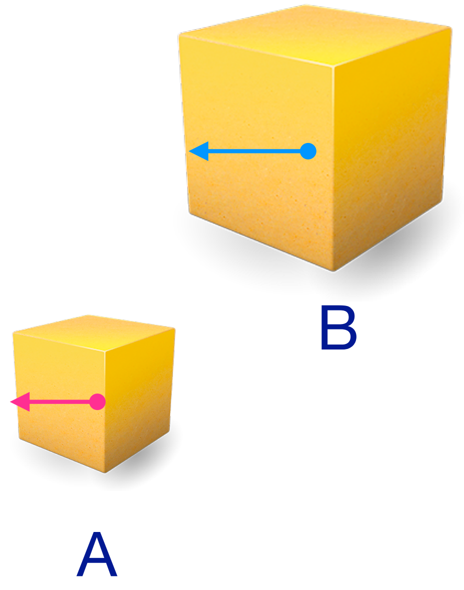
- Energy Transfer/Change
  - by Heat
    - Heat (Temperature Change)
    - Heat Capacity
      - Specific Heat Capacity
      - Molar Heat Capacity
    - Heat Transfer
  - by Work
    - Pressure-Volume Work
  - Measuring Changes in Internal Energy
    - Bomb Calorimetry

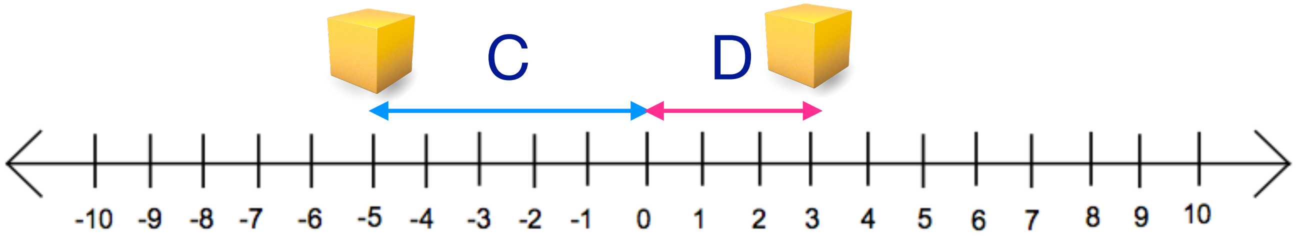




Measurements of Size & Position

- Properties can be intensive or extensive.
  - Measurements of extensive properties answer the question “how much?”
    - These are measurements of size (extent).
    - They say how many units are contained in the sample.
    - Example:
      - Mass, contains 50.0 grams
      - Volume, contains 32 mL
  - Measurements of intensive properties answer the question “how far?”
    - These are measurements of position.
    - They say how many units are between that value and a reference point.
      - Measurements of position can be positive or negative.
      - They require a reference point.
    - Example:
      - Speed, 55 mph faster (than another object)
      - Brightness, 0.23 lumen brighter (than another source)
      - Hardness, 7.0 Moh (70% the hardness of diamond)
      - Density, 19.3 g/mL (19.3 times as dense as water)





Change (Δ)

- Change is the difference between two measurements, over time.
  - We always assume time moves forward.
  - So change is always final minus initial.
- The change of a value is represented with the delta symbol (Δ).

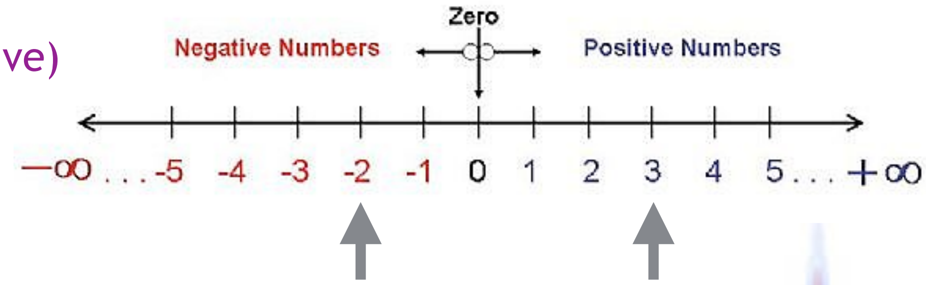
$$\Delta X = X_{final} - X_{initial}$$

X_{FINAL} is -2.

$X_{INITIAL}$ is +3.

- Change can be represented by:
  - A single value (which can be positive or negative)

$\Delta X = -5$



X_{FINAL} $X_{INITIAL}$

  - A size and a direction (example: decreased)

X decreased by +5

(the size of the change is +5)

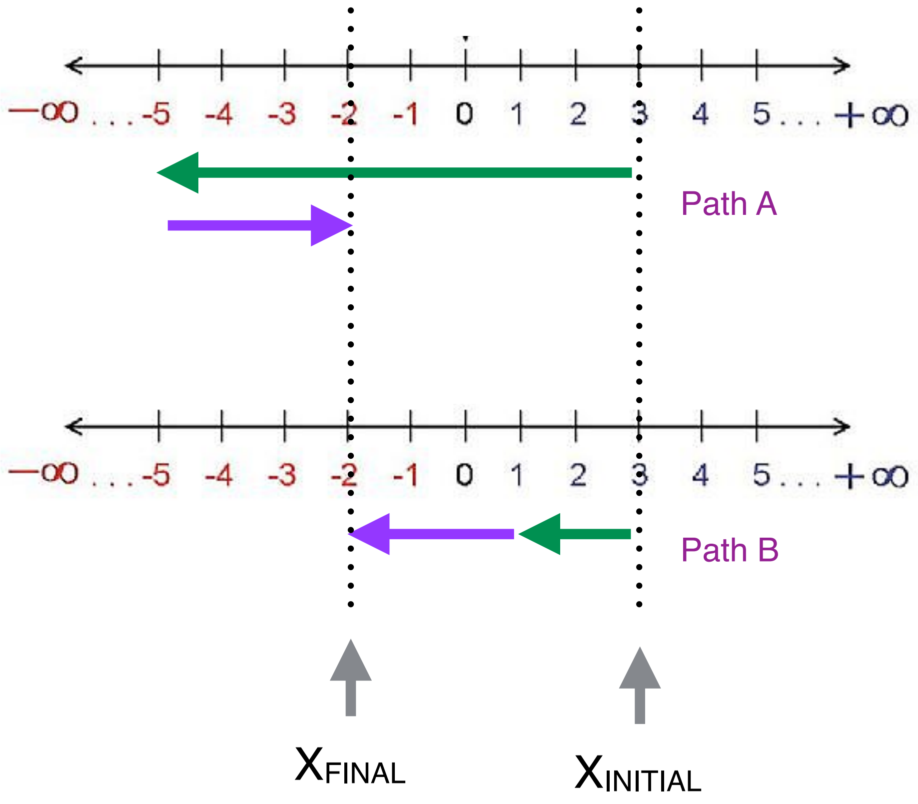
The change in X is

(-2) - (+3) = -5

$\Delta X = -5$

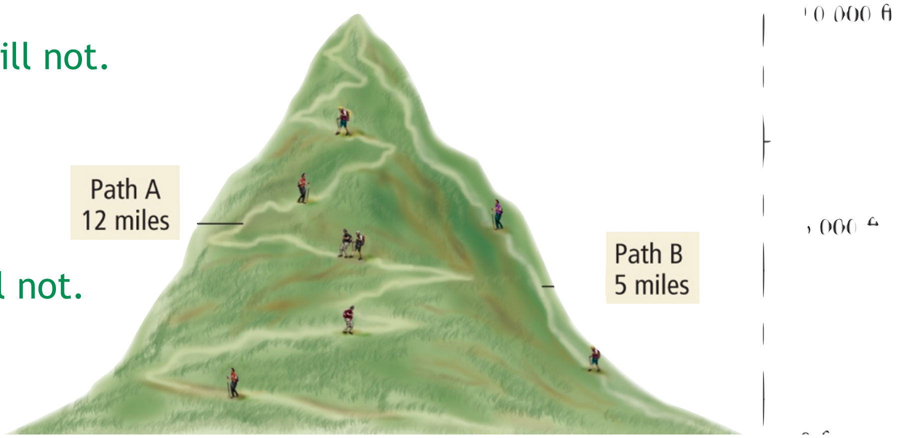
Measuring Changes

- A path is a group of steps you took to accomplish a change.
  - Different paths can be used to change from the same initial state to the same final state.
  - You can go from +3 to -2 by:
    - subtracting 8 and adding 3
    - subtracting 2 and subtracting 3
    - or many other paths
- Some properties associated with change depend on the path you took to produce the change, others don't.
  - If I drive from Madison WI to Helena MT.
    - … how much gasoline I use will be a function of my path.
    - … my change in time zones will not.
  - If I climb a hill
    - … the miles I walk will be a function of my path
    - … my change in elevation will not.



Measuring Changes

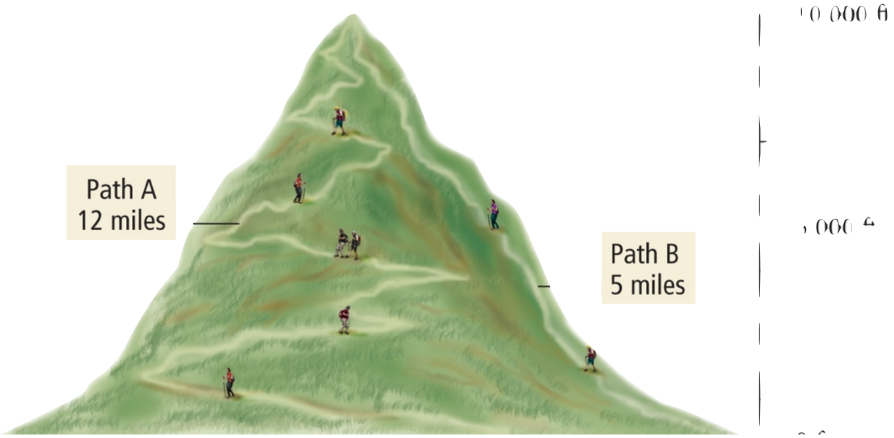
- A path is a group of steps you took to accomplish a change.
  - Different paths can be used to change from the same initial state to the same final state.
  - You can go from +3 to -2 by:
    - subtracting 8 and adding 3
    - subtracting 2 and subtracting 3
    - or many other paths
- Some properties associated with change depend on the path you took to produce the change, others don't.
  - If I drive from Madison WI to Helena MT.
    - … how much gasoline I use will be a function of my path.
    - … my change in time zones will not.
  - If I climb a hill
    - … the miles I walk will be a function of my path
    - … my change in elevation will not.



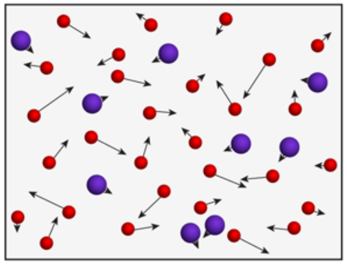
Measuring Changes

- A state property is a property that can be measured considering only a systems condition — it’s state.
  - A state function is the equation that calculates the property.
- A path property is a property that requires you to measure it’s path to determine the property.
  - A path function is the equation that calculates the property.

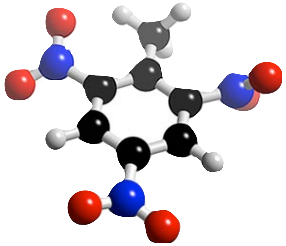
- Example 1:
  - The fuel level in a gas tank is a state property.
  - How much fuel went into the tank is a path property..
- Example 2:
  - Altitude can be calculated with a state function.
  - Miles traveled must be calculated with a path function



Internal Energy



- A system is any portion of the universe we single out for study.
- Internal Energy is the energy contained within that system.
  - Chemists care about:
    - the thermal energy (kinetic energy)
    - the chemical energy (potential energy)



- Work (w) is the energy used to cause an object to move against a force.
    - Lifting a box, pulling a bow, throwing a stone.
- Heat (q or Q) is energy that flows between two samples due to their difference in temperature.
    - Thermal energy moving from a hot material to a cold one.

$$\Delta E_{sys} = q + w$$

- Internal Energy is a state function. (*this will be very important*)
- Work and Heat are path functions.

- We can't easily measure internal energy. But we can measure work and heat. If we know the work and heat that is done to a system, we can know how internal energy changes.

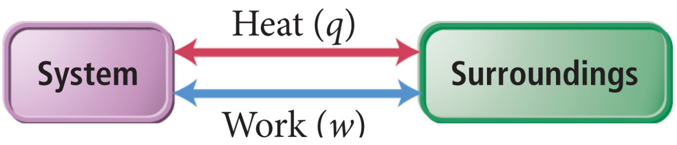
Why it matters that ΔE_{sys} is a state function.

- Internal energy is all chemical energy and thermal energy in a system.
- How much energy is in a burning coal, an ice cube, or 4,000 L of petroleum is useful information.
- We have no idea how to measure internal energy. (practically)
- However it is possible to measure work and heat.
  - And if we add all the work and heat related to a process we can calculate the change in internal energy of that process.
- We can do an experiment to find out how the internal energy changes when a coal burns, ice cube melts, or 4,000 L of petroleum are consumed.
- And because ΔE_{sys} is a state function once we know it's value for that coal, ice cube or 4,000 L of petroleum we know it for every identical lump of coal, cube of ice, or equivalent volume of petroleum.

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

$$\Delta E = E_{products} - E_{reactants}$$

$$\Delta E_{sys} = q + w$$



Finding ΔE_{sys}

$$\Delta E_{sys} = q + w$$

TABLE 5.1 ■ Sign Conventions for q, w, and ΔE

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

- When we use q and w it's **extremely important** we keep our perspective consistent so use the correct sign with each number!
- Example: A system undergoes the following processes, what is the net change in internal energy?
  - The system releases 15 J heat while doing 65 J of work.

$$\Delta E = q + w = -15\text{ J} + -65\text{ J} = -80\text{ J}$$

  - The system does 34 J of work while absorbing 67 J of heat.

$$\Delta E = q + w = +67\text{ J} + -34\text{ J} = 33\text{ J}$$

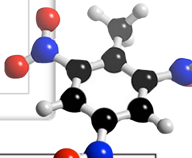
  - The system absorbs 35 J of heat, releases 15 J of heat, and has 23 J of work done on it.

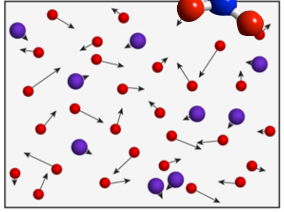
$$\Delta E = q + w = +35\text{ J} + -15\text{ J} + 23\text{ J} = 43\text{ J}$$

So how do we find q?

Ch06

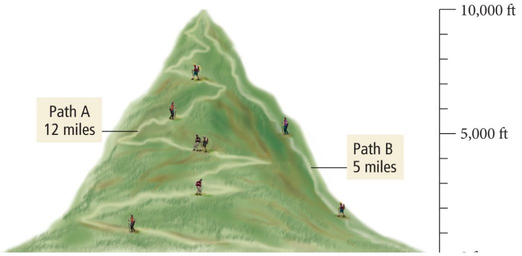
Accounting for Energy

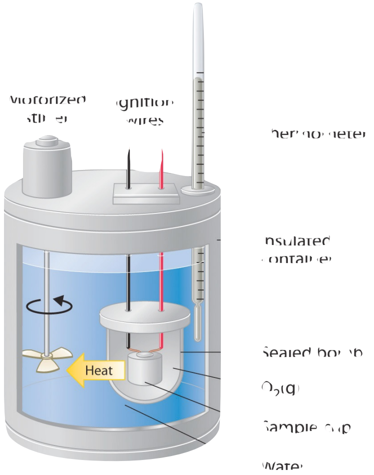




- Energy
  - Definitions
  - Classifications
    - Kinetic, Potential, Thermal & Chemical
  - Units
- Systems – keeping perspective consistent is essential!
  - the system & the surroundings
  - types of systems
    - open, closed, isolated
- Changes, Values & Size
  - What is change? (over time)
  - Measuring changes
    - State Functions vs Path Functions
  - Changes in Energy
    - Internal Energy
    - Heat & Work

- Energy Transfer/Change
  - by Heat
    - Heat (Temperature Change)
    - Heat Capacity
      - Specific Heat Capacity
      - Molar Heat Capacity
    - Heat Transfer
  - by Work
    - Pressure-Volume Work
  - Measuring Changes in Internal Energy
    - Bomb Calorimetry

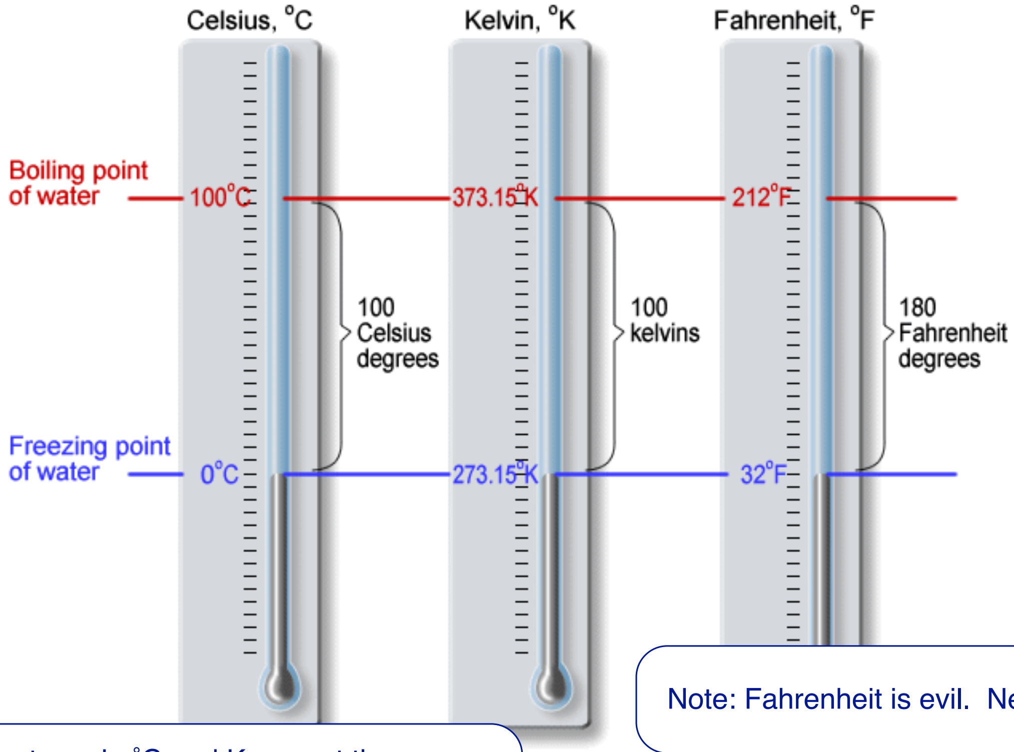




A word about ΔT…

ΔT in Kelvin or Celsius is the Same Number

- A temperature **T** (value) is not the same in Celsius, Kelvin or Fahrenheit.
  - The size of each degree is the same in Celsius and Kelvin.
- A change in temperature **ΔT** (change) is the same in Celsius and Kelvin.



We generally measure temperature in Celsius, it must be in Kelvin for most calculations.

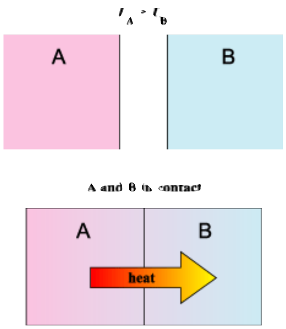
For equations involving ΔT this conversion is trivial and often skipped.

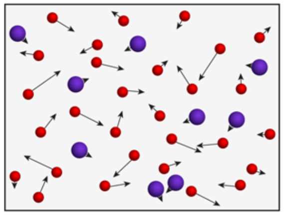
Note: Fahrenheit is evil. Never use it. For anything.

Note: Temperatures in ˚C and K are not the same.
But a difference in ˚C and K are identical!

Introducing Heat Capacity

- Heat (q or Q) is energy that flows between two systems due to their difference in temperature.
  - Observation: The more energy an object gains, the hotter it becomes (it's temperature goes up).
  - Observation: Different objects have different capacities to hold heat, before their temperature increases.

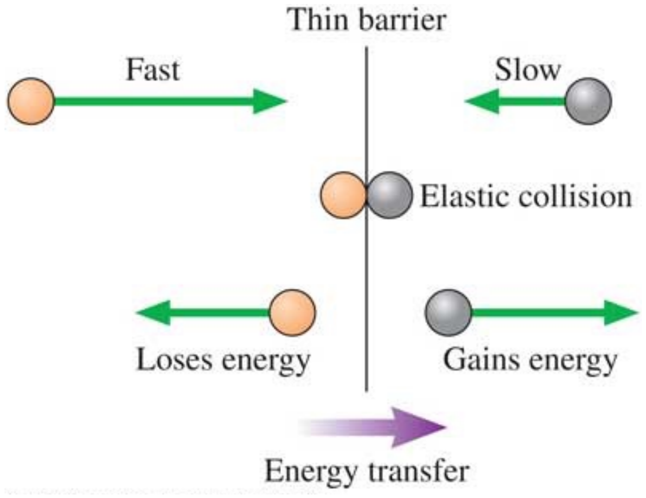


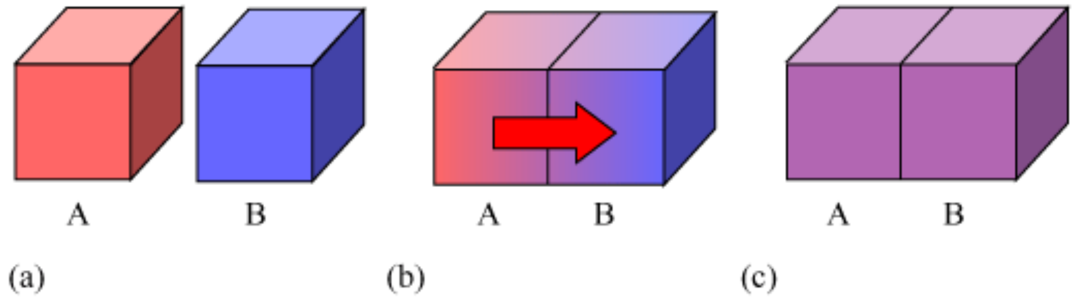


↑E ↑T

$\Delta E \propto \Delta T$

$q \propto \Delta T$



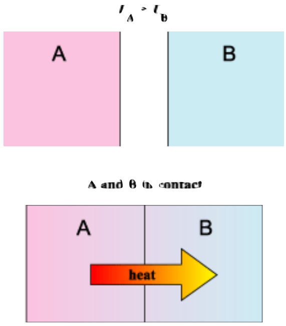


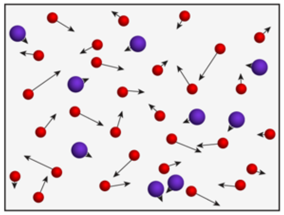
Introducing Heat Capacity

- Heat (q or Q) is energy that flows between two systems due to their difference in temperature.
    - Observation: The more energy an object gains, the hotter it becomes (it's temperature goes up).
    - Observation: Different objects have different capacities to hold heat, before their temperature increases.

- Heat capacity (C) is the heat an object can hold before it rises 1˚. Think of this as resistance to heat or capacity to hold heat at each degree.
    - Heat Capacity varies between objects.
        - Size matters, a bigger frying pan can hold more heat than a smaller one.
        - Substance matters, an oven mitt takes more heat to raise it's temperature than the frying pan.

- If we want to explore heat capacity as a function of the substance we need to factor out size (amount of substance).





$$\uparrow E \qquad \uparrow T$$

$$\Delta E \;\alpha\; \Delta T$$

$$q \;\alpha\; \Delta T$$

$$\boxed{q = C \cdot \Delta T}$$

Specific Heat Capacity & Molar Heat Capacity

- **Heat capacity** (C) is the heat an object can hold before it rises 1˚. Think of it as resistance to heat or capacity to hold heat at each degree.
  - We can factor out size (amount of substance), by dividing heat capacity by the moles or grams of a substance.
- This gives us two new kinds of heat capacity:
  - **Molar Heat Capacity** (C_m) is the heat capacity per mole (n) of a substance.

    $C_m = C / n$ or $C = C_m * n$

  - **Specific Heat Capacity** (C_s) is the heat capacity per gram (m) of a substance.
    (Often abbreviated as Specific Heat)

    $C_s = C / m$ or $C = C_s * m$

- These two new heat capacities are a property of the substance! They don't change!
- You might need to use any of the three heat capacity equations depending on the problem.

There are three forms of the heat capacity equation!

$$q = C \cdot \Delta T$$
$$q = C_m \cdot n \cdot \Delta T$$
$$q = C_s \cdot m \cdot \Delta T$$

Specific Heat and Molar Heat Capacity are related by molar mass.

$$C_m \cdot n = C_s \cdot m$$
$$C_m / C_s = m/n$$
$$= \text{grams/moles}$$
$$= \text{molar mass}$$

C = 2,835 J / ˚C

C_s = 0.45 J / g ˚C

C_m = 25.1 J / mol ˚C

C = 62.0 J / ˚C

Note: Temperatures in ˚C and K are not the same.
But a difference in ˚C and K are identical!

Heat Capacity

- A 12.5 gram block of metal cools from 53.2˚C to 28.7˚C. It's specific heat is 19 J/g-K. How much energy was released?

The system is our block

$\Delta T = T_F - T_I$
$= 28.7°C - 53.2°C$
$= -24.5°C$

$q = C_s \cdot m \cdot \Delta T$
$= 19 J/gK \cdot 12.5g \cdot (-24.5°C)$
$= -5818.75 J$
$= -5.8 kJ$

28.7
− 53.2
− 24.5

The energy in the system went down by 5.8 kJ

$\boxed{\text{+5.8 kJ were released from the block.}}$

$q = C_s \cdot m \cdot \Delta T$

Problem:

How much heat does it take to warm 143 grams of water from 25.0˚C to 78.0˚C. The specific heat of water is 4.184 J/g K.

Solution

Calculate ΔT ; Apply Specific Heat Eqn.

$q = C_s \cdot m \cdot \Delta T$

$C_s = 4.184\ J/g\ K$

$T_F = 78.0°C$

$T_I = 25.0°C$

$m = 143\ g$

Part A

$\Delta T = T_F - T_I$

$= 53.0°C$

$= 53.0\ K$

$$\begin{array}{r} 78.0 \\ -\ 25.0 \\ \hline 53.0 \end{array}$$

Part B

$q = C_s \cdot m \cdot \Delta T$

$= 4.184\ J/g°K \cdot 143 g \cdot 53.0 K$

$= 3.17 \times 10^4\ J$

$\boxed{\checkmark = 31.7\ kJ}$

Problem:

A 62.0˚C iron cube weighing 149 grams looses 2.34 kJ of energy, what is it’s final temperature? The specific heat of iron is 0.450 J/g K.

Solution

$$q = -2.34\ \text{kJ}$$
$$= -2.34 \times 10^3\ \text{J}$$

$$q = C_s \cdot m \cdot \Delta T$$
$$C_s = 0.450\ \text{J/gK}$$
$$m = 149\ \text{g}$$
$$T_I = 62.0\ ^\circ\text{C}$$
$$\Delta T = T_F - T_I$$
$$T_F = ?$$

Calculate ΔT ; Solve for T_{FINAL}

Part A

$$q = C_s \cdot m \cdot \Delta T$$
$$\Delta T = \frac{q}{C_s \cdot m}$$
$$= \frac{-2.34 \times 10^3\ \text{J}}{0.450\ \text{J/gK} \cdot 149\ \text{g}}$$
$$= -34.9\ \text{K}$$

for a change in T

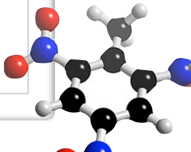
$$\Delta T = -34.9\ ^\circ\text{C}$$

Part B

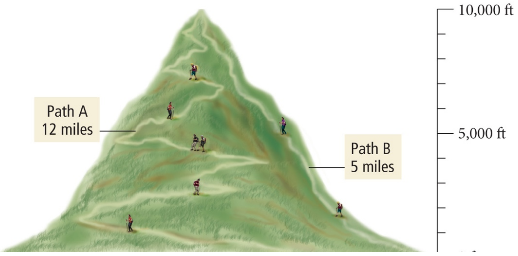
$$\Delta T = T_F - T_I$$
$$T_F = \Delta T + T_I$$
$$= -34.9\ ^\circ\text{C} + 62.0\ ^\circ\text{C}$$
$$= 27.1\ ^\circ\text{C}$$

$$\boxed{27.1\ ^\circ\text{C}}$$

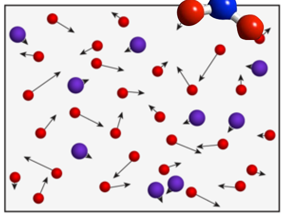
Accounting for Energy

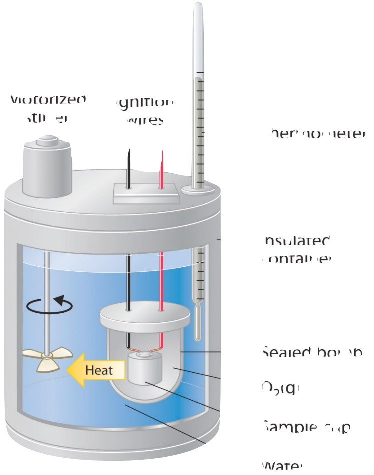


- Energy
  - Definitions
  - Classifications
    - Kinetic, Potential, Thermal & Chemical
  - Units
- Systems – keeping perspective consistent is essential!
  - the system & the surroundings
  - types of systems
    - open, closed, isolated
- Changes, Values & Size
  - What is change? (over time)
  - Measuring changes
    - State Functions vs Path Functions
  - Changes in Energy
    - Internal Energy
    - Heat & Work



- Energy Transfer/Change
  - by Heat
    - Heat (Temperature Change)
    - Heat Capacity
      - Specific Heat Capacity
      - Molar Heat Capacity
    - Heat Transfer
  - by Work
    - Pressure-Volume Work
  - Measuring Changes in Internal Energy
    - Bomb Calorimetry





Energy Transfer

The law of conservation of matter:

Matter can never be created nor destroyed.

The first law of thermodynamics:

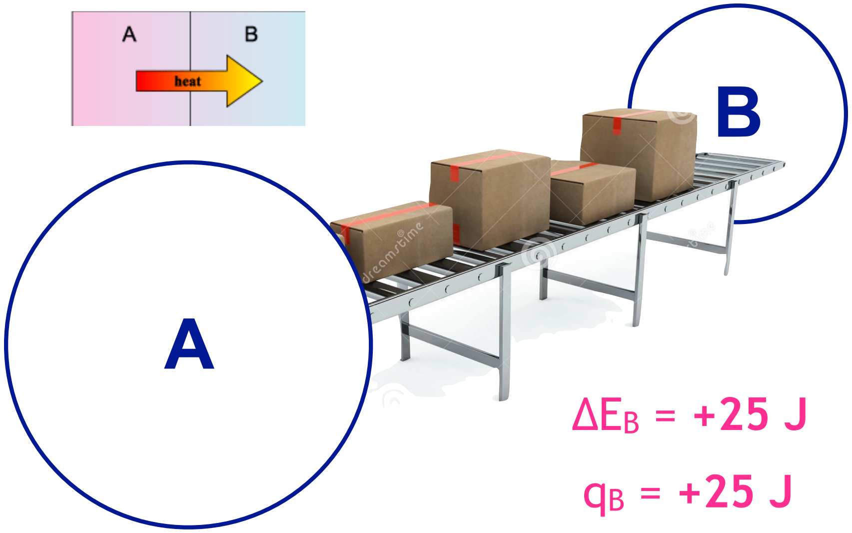
Energy can never be created nor destroyed.

At the atomic scale and above, there are no exceptions to these laws.

- So energy that leaves one system must go into another.
  - If one system has a negative ΔE another must have a positive ΔE
  - If we consider only heat, that means if one system has a positive q another must have a negative q.
- If A Looses 25 J then then B gains 25 J
  - The size of both changes will be 25 J
  - The direction of the changes will be opposite.

$$q_A = -1 \times q_B$$

- Note words like gain or loose provide that direction so be careful to not use a double negative:
- Never say q_b lost -25 J!



ΔE_B = **+25 J**

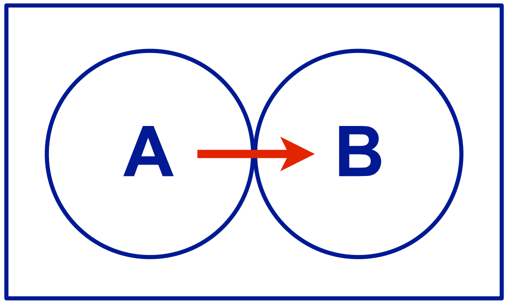
q_B = **+25 J**

B gains **+25 J**

ΔE_A = **-25 J**

q_A = **-25 J**

A looses **+25 J**



Note: Temperatures in ˚C and K are not the same.
But a difference in ˚C and K are identical!

Heat Capacity / Heat Transfer

- A 12.5 gram block of metal cools from 53.2˚C to 28.7˚C. It's specific heat is 19 J/g-K. How much energy was released?

The system is our block

$\Delta T = T_F - T_I$

$= 28.7°C - 53.2°C$

$= -24.5\ °C$

$28.7 - 53.2 = -24.5$

$q = C_s \cdot m \cdot \Delta T$

$= 19\ J/gK \cdot 12.5\ g \cdot (-24.5°C)$

$= -5818.75\ J$

$= -5.8\ kJ$

The energy in the system went down by 5.8 kJ

$q = C_s \cdot m \cdot \Delta T$

$\boxed{\text{+5.8 kJ were released from the block.}}$

- It cooled in water. How much energy was gained by the water?

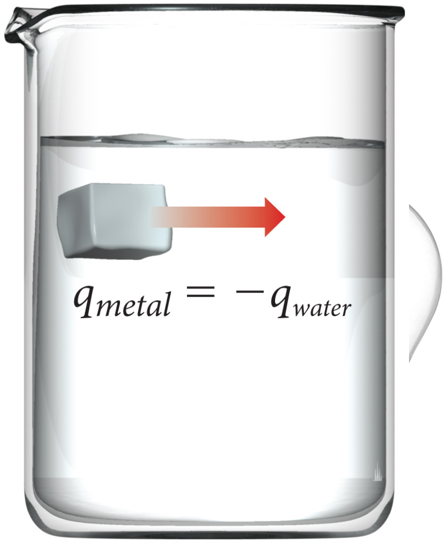
$q_{METAL} = -q_{WATER}$

$-5.8\ kJ = -q_{WATER}$

$q_{WATER} = +5.8\ kJ$

$\Delta E_{METAL} = -5.8\ kJ$

$\Delta E_{WATER} = +5.8\ kJ$



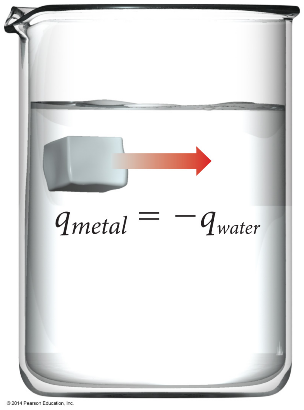
© 2014 Pearson Education, Inc.

The block lost 5.8 kJ

$\boxed{\text{The water gained 5.8 kJ}}$

Heat Capacity

- A 12.5 gram block at 53.2˚C is dropped into 255 grams of liquid at 25.8˚C. The specific heat of the metal is 19 J/g-K. The liquid and metal come to a final temperature of 32.7˚C. What is the heat capacity of the liquid?



① Find ΔT of Metal

② Find q of metal

③ Find q of liquid

④ Find Cs of liquid

METAL

$C_s = 19\ J/gK$

$m = 12.5\,g$

$T_I = 53.2°C$

$T_F = 32.7°C$

Liquid

$m = 255\,g$

$T_I = 25.8°C$

$T_F = 32.7°C$

$$\Delta T_{METAL} = T_F - T_I$$
$$= 32.7°C - 53.2°C$$
$$= -20.5°C$$

$$\begin{array}{r|l} 32.7 & °C \\ -\ 53.2 & °C \\ \hline -20.5 & °C \end{array}$$

$$q_{METAL} = C_s \cdot m \cdot \Delta T$$
$$= 19\,J/gK \cdot 12.5\,g \cdot (-20.5°C)$$
$$= -4868.75\ J = -4800\ J$$

$$q_{METAL} = -q_{WATER}$$
$$-4800\ J = -q_{LIQ}$$
$$4800\ J = q_{LIQ}$$

$$\Delta T_{LIQ} = T_{F_{LIQ}} - T_{I_{LIQ}}$$
$$= 32.7°C - 25.8°C$$
$$= 6.9\ K$$

$$\begin{array}{r|l} 32.7 & °C \\ -\ 25.8 & °C \\ \hline 6.9 & °C \end{array}$$

$$q_{LIQ} = C_{s_{LIQ}} \cdot m_{LIQ} \cdot \Delta T_{LIQ}$$
$$4800\ J = C_{s_{LIQ}} \cdot 255\,g \cdot 6.9°C$$
$$C_{s_{LIQ}} = \frac{4800\ J}{255\,g \cdot 6.9\,K}$$
$$= \boxed{2.7\ J/gK}$$

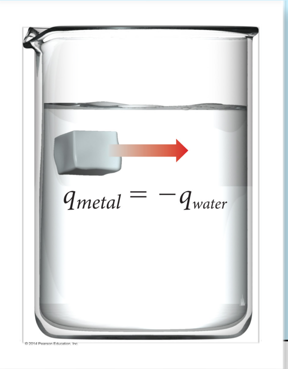
Problem: Metal Cooling in Water

A 12.48 g sample of an unknown metal is heated to 99.0 °C and then was dropped into 50.0 grams of 25.0 °C water. The temperature of the water rose to 28.1°C, water has specific heat of 4.184 J/g-K.

(a) How many joules of energy did the water absorb?

(b) How many joules of energy did the metal lose?

(c) What is the specific heat capacity of the metal?



Metal

$m = 12.48 g$

$T_I = 99.0 °C$

$T_F = 28.1 °C$

$C_s = ?$

$q_m = ?$

WATER

$m = 50.0 g$

$T_I = 25.0 °C$

$T_F = 28.1 °C$

$C_s = 4.184 J/gK$

$q_w = ?$

① Find q_{WATER} ② Find q_{METAL} ③ Find C_s metal

WATER

$\Delta T = T_F - T_I$

$$\begin{array}{r} 28.1\,°C \\ -\;25.0\,°C \\ \hline 3.1\,°C \end{array}$$

$\Delta T = 3.1 K$

(Note: $\Delta °C = \Delta K$)

$q_w = C_s \cdot m \cdot \Delta T$

$= (4.184 J/gK)(50.0 g)(3.1 °C)$

$= 650 J$

Metal

$q_M = -q_w = -650 J$

$\Delta T = T_F - T_I$

$= -70.9 K$

$$\begin{array}{r} 28.1\,°C \\ -\;99.0\,°C \\ \hline -70.9\,°C \end{array}$$

$q_m = C_s \cdot m \cdot \Delta T$

$C_s = \frac{q_m}{m \cdot \Delta T} = \frac{-650 J}{12.48 g \cdot (-70.9 K)}$

$= 7.3 J/gK$

(a) 650 J

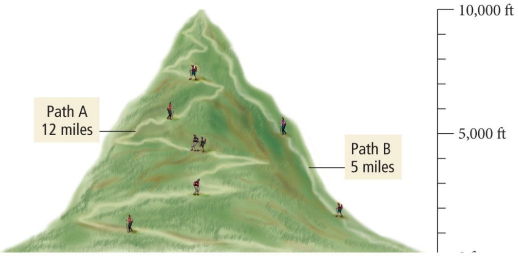
(b) 650 J

(c) 7.3 J/gK

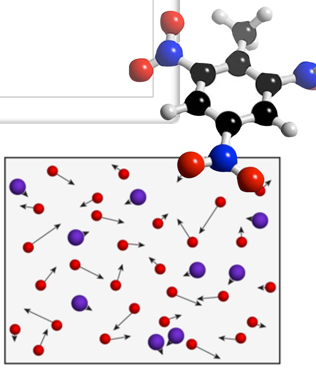
Ch06

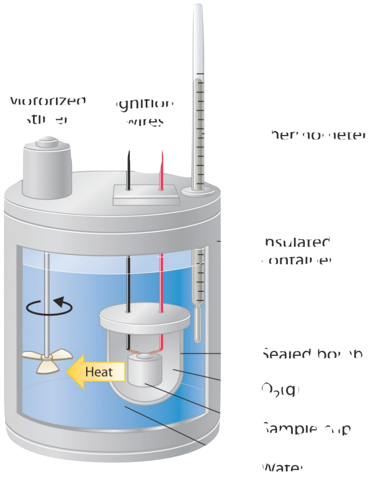
Accounting for Energy

- Energy
  - Definitions
  - Classifications
    - Kinetic, Potential, Thermal & Chemical
  - Units
- Systems – keeping perspective consistent is essential!
  - the system & the surroundings
  - types of systems
    - open, closed, isolated
- Changes, Values & Size
  - What is change? (over time)
  - Measuring changes
    - State Functions vs Path Functions
  - Changes in Energy
    - Internal Energy
    - Heat & Work

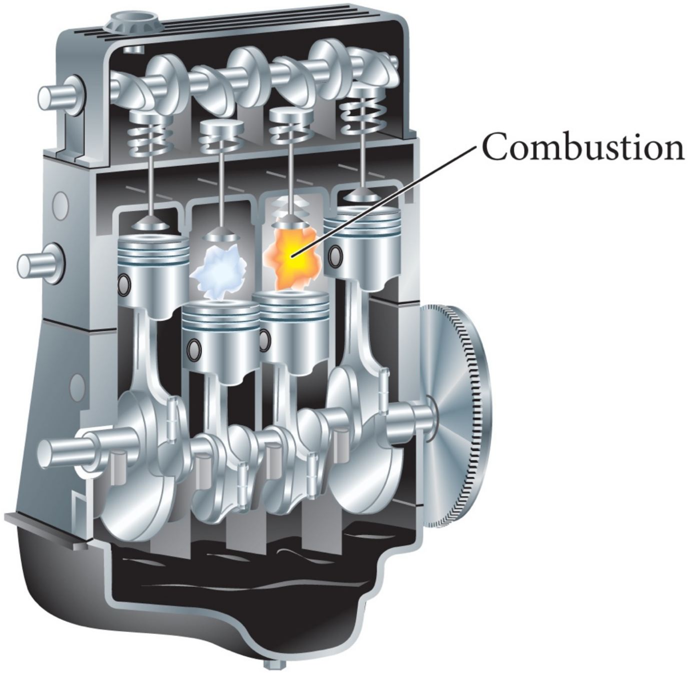


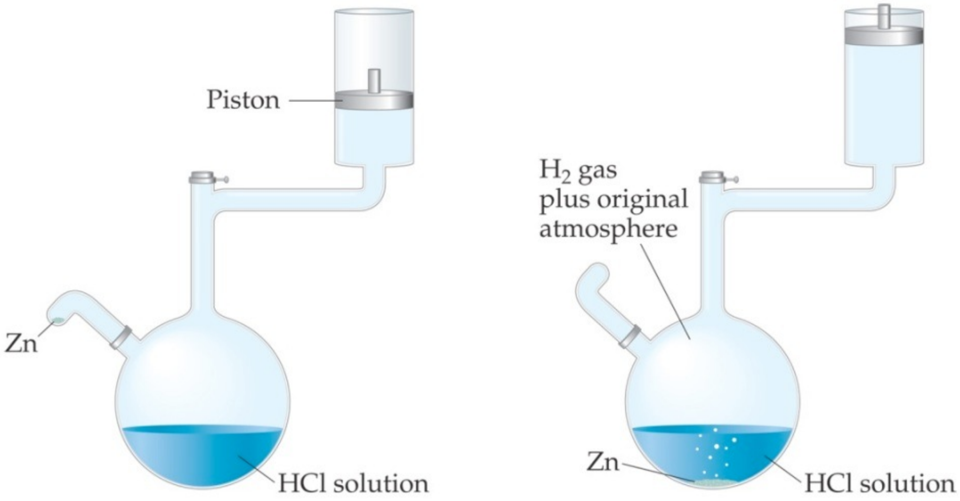
- Energy Transfer/Change
  - by Heat
    - Heat (Temperature Change)
    - Heat Capacity
      - Specific Heat Capacity
      - Molar Heat Capacity
    - Heat Transfer
  - by Work
    - Pressure-Volume Work
  - Measuring Changes in Internal Energy
    - Bomb Calorimetry





Pressure-Volume Work





- Heat (q) is one way to change internal energy.
- Another way is work (w). Work is the energy used to move an object.
- Heat and work are the primary ways the internal energy of chemical systems change.

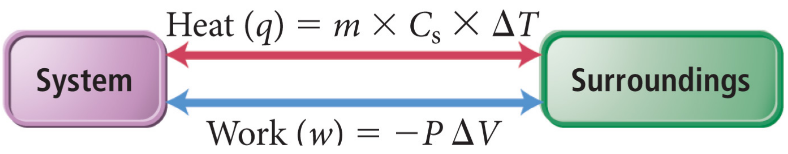
$$\Delta E = q + w$$

- Chemical systems do work by changing volume under pressure.

$$P \times \Delta V$$

- For example, burning gasoline converts it to gases which pushes a piston.
- By convention, w and ΔV must be opposite in sign, so w and ΔE can have the same sign

$$w = -P \times \Delta V$$



$\text{Heat } (q) = m \times C_s \times \Delta T$

$\text{Work } (w) = -P\,\Delta V$

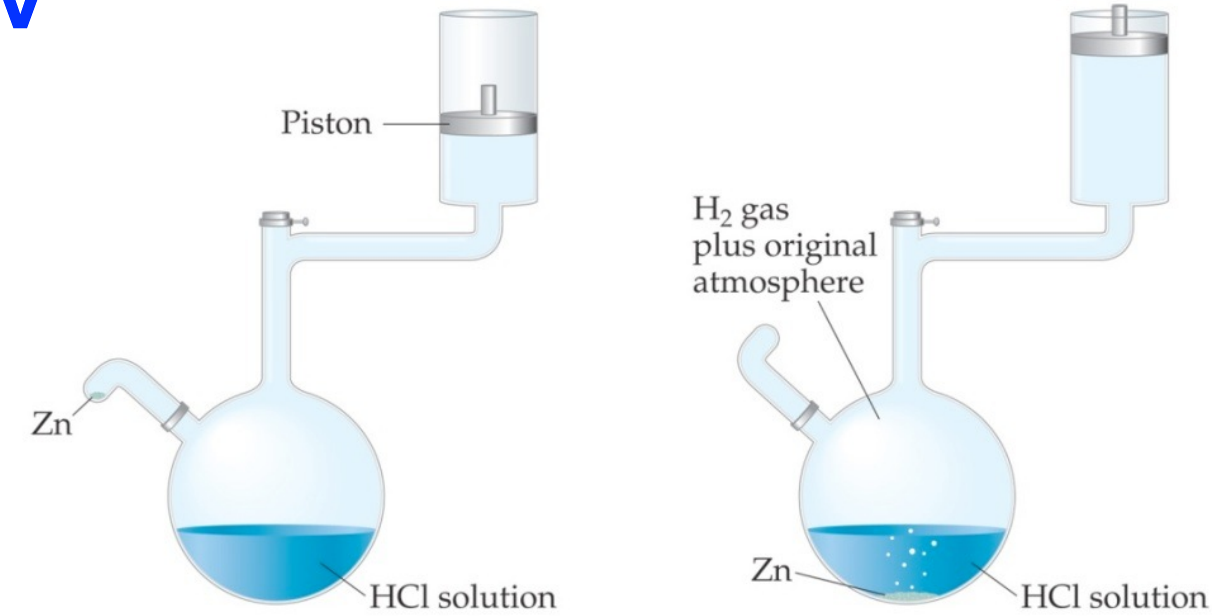
$$\Delta E = q - P \times \Delta V$$

Pressure-Volume Work

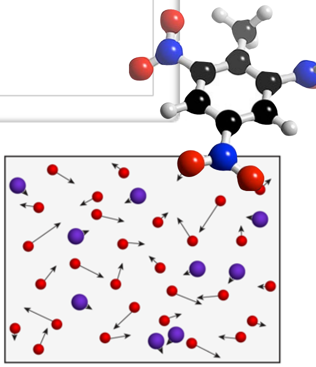
Chemical reactions most commonly work on their surroundings (“push things”) by changing volume and thereby exerting pressure.

$$w = -P \Delta V$$

(reminder: positive work is work done on the system, so work done by the system is negative)

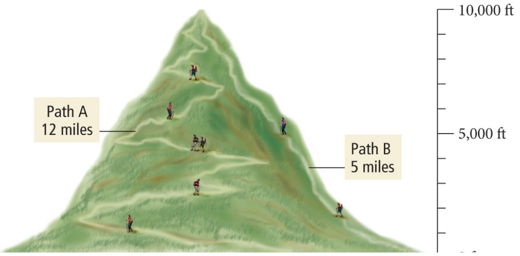


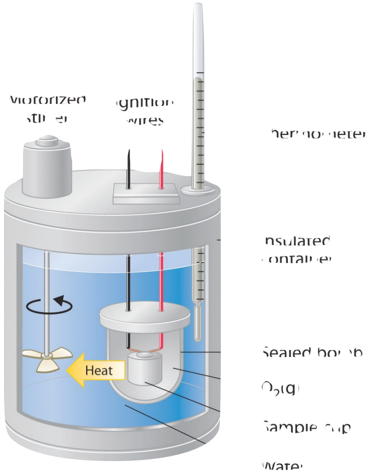
Accounting for Energy



- Energy
  - Definitions
  - Classifications
    - Kinetic, Potential, Thermal & Chemical
  - Units
- Systems – keeping perspective consistent is essential!
  - the system & the surroundings
  - types of systems
    - open, closed, isolated
- Changes, Values & Size
  - What is change? (over time)
  - Measuring changes
    - State Functions vs Path Functions
  - Changes in Energy
    - Internal Energy
    - Heat & Work

- Energy Transfer/Change
  - by Heat
    - Heat (Temperature Change)
    - Heat Capacity
      - Specific Heat Capacity
      - Molar Heat Capacity
    - Heat Transfer
  - by Work
    - Pressure-Volume Work
  - Measuring Changes in Internal Energy
    - Bomb Calorimetry

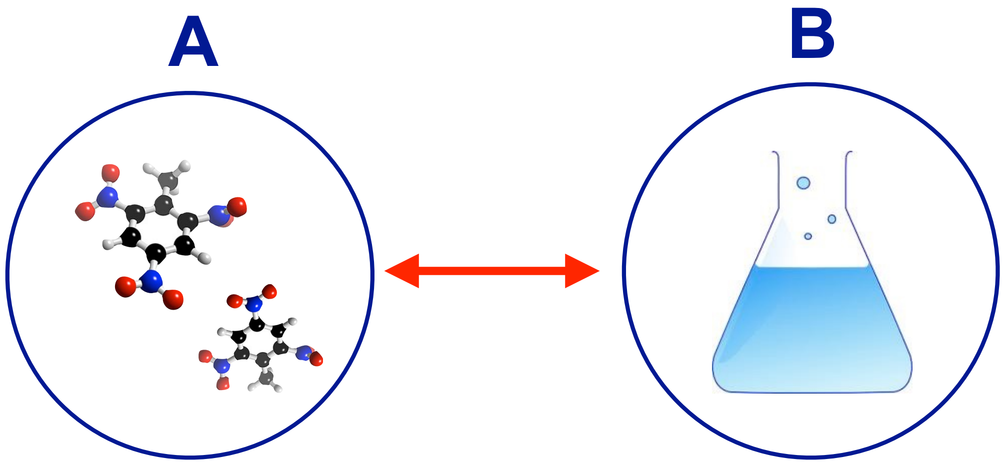




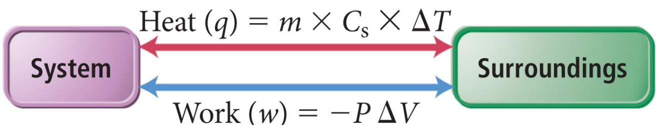
Calorimetry

- We’re interested in chemical reactions.
- We want to know how the energies change when one substance turns into another.
- We want to know if we’re putting energy into a substance by chemical reaction, or if we’re releasing it.
- We know that energy is the result of all work and heat that occurs in the reaction.
- Heat can be determined by measuring temperature and knowing heat capacity.
- But work is a lot harder to track.
- So how do we find that change?

- Measuring energy changes is called calorimetry.



$C_3H_8 + 8\ O_{2\ (g)} \rightarrow 3\ CO_{2\ (g)} + 4\ H_2O_{\ (g)}$


Heat $(q) = m \times C_s \times \Delta T$

System

Surroundings

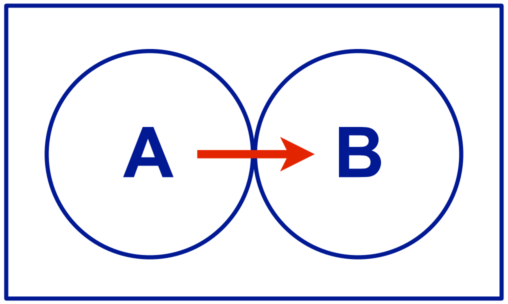
Work $(w) = -P\,\Delta V$

$$\Delta X = X_{final} - X_{initial}$$

$$\Delta E = E_{Products} - E_{Reactants}$$

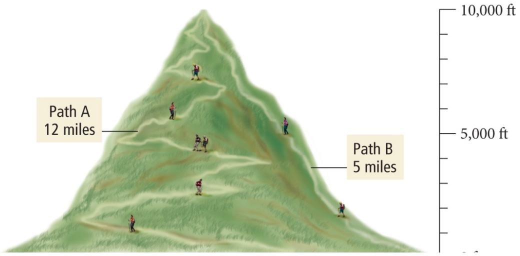
$$\Delta E = q + w$$

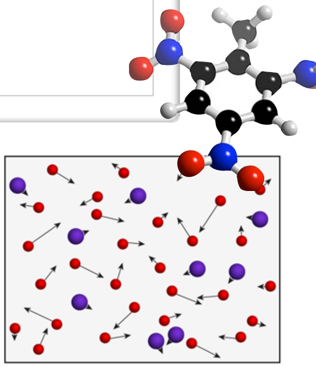
$$E_{Products} - E_{Reactants} = q + w$$



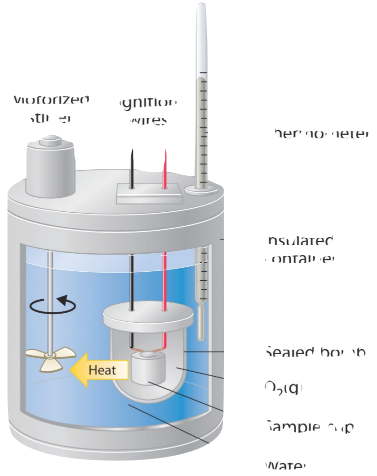
Accounting for Energy

- Energy
  - Definitions
  - Classifications
    - Kinetic, Potential, Thermal & Chemical
  - Units
- Systems – keeping perspective consistent is essential!
  - the system & the surroundings
  - types of systems
    - open, closed, isolated
- Changes, Values & Size
  - What is change? (over time)
  - Measuring changes
    - State Functions vs Path Functions
  - Changes in Energy
    - Internal Energy
    - Heat & Work





- Energy Transfer/Change
  - by Heat
    - Heat (Temperature Change)
    - Heat Capacity
      - Specific Heat Capacity
      - Molar Heat Capacity
    - Heat Transfer
  - by Work
    - Pressure-Volume Work
  - Measuring Changes in Internal Energy
    - Bomb Calorimetry



Bomb Calorimetry

$$\Delta E = q + w$$

$$w = -P\,\Delta V$$

$$\Delta E = q + (-P\,\Delta V)$$

if constant volume: then $\Delta V = 0$

$$\Delta E = q_v + (-P)(0)$$

$$\Delta E = q_v$$

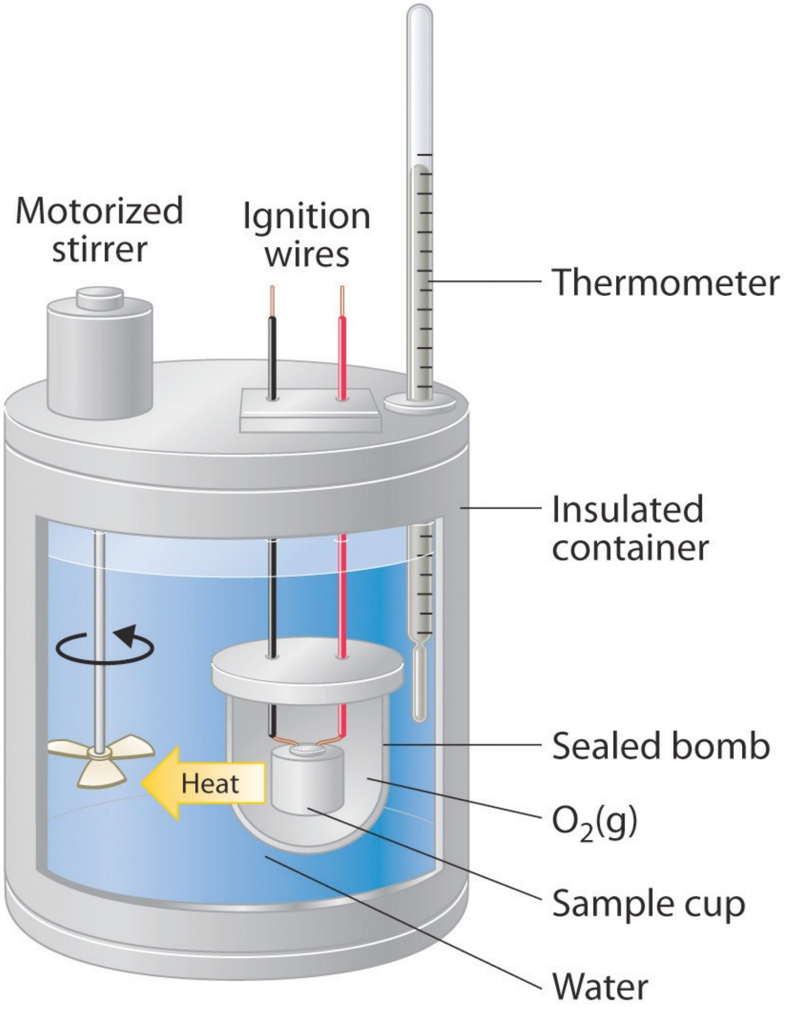
This means, if volume is constant we can measure changes in internal energy just by measuring heat flow q_v

$$q = C \cdot \Delta T$$

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

$$q = C_s \cdot m \cdot \Delta T$$

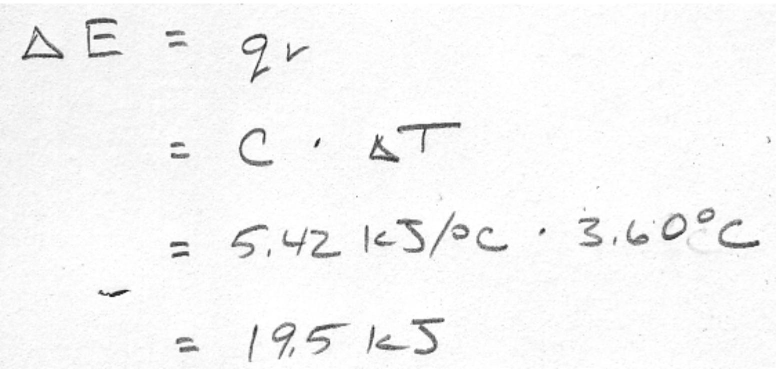
A bomb is a rigid sealed container for chemical reactions.



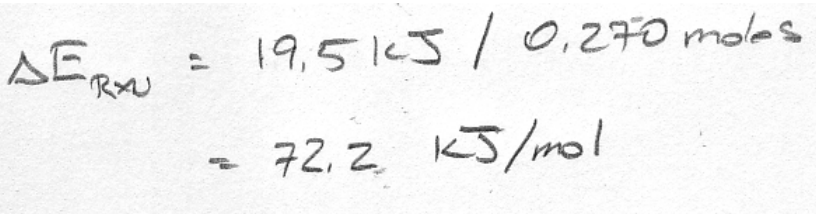
Finding ΔE w/ Bomb Calorimetry

- A bomb calorimeter has a heat capacity of 5.42 kJ/˚C. If you burn 0.270 moles in the calorimeter and see the temperature rise 3.60˚C, what is the reactions ΔE in kJ/mole?

$$q = C \cdot \Delta T$$
$$q = C_m \cdot n \cdot \Delta T$$
$$q = C_s \cdot m \cdot \Delta T$$

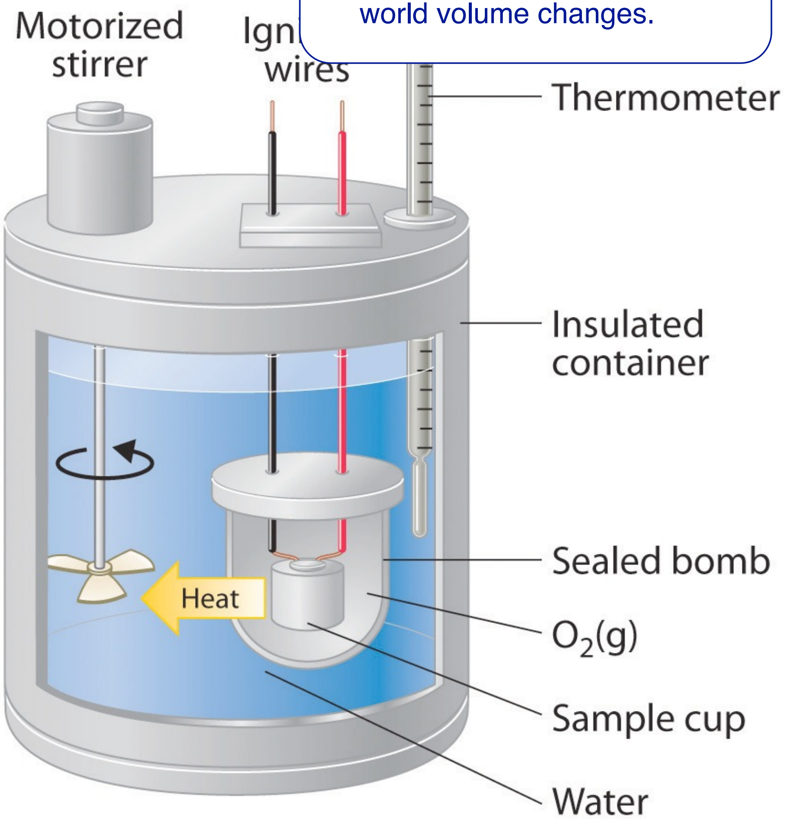


$$\Delta E = q_v$$
$$= C \cdot \Delta T$$
$$= 5.42\ \text{kJ/°C} \cdot 3.60\ \text{°C}$$
$$= 19.5\ \text{kJ}$$



$$\Delta E_{Rxn} = 19.5\ \text{kJ} / 0.270\ \text{moles}$$
$$= 72.2\ \text{kJ/mol}$$

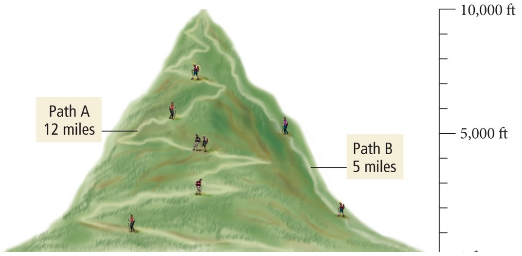
Disadvantages:
- the equipment is expensive
- the experiment is difficult
- it’s not exactly the reaction we’re interested in, in the real world volume changes.



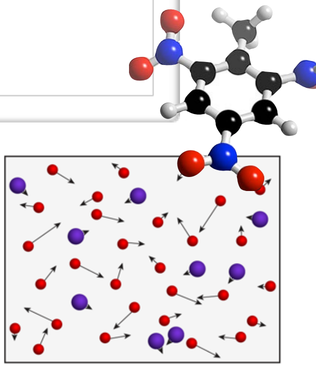
Ch06

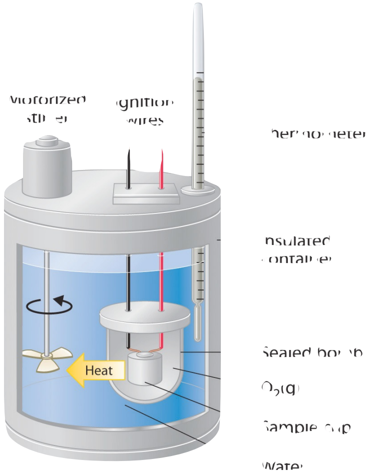
Accounting for Energy

- Energy
  - Definitions
  - Classifications
    - Kinetic, Potential, Thermal & Chemical
  - Units
- Systems – keeping perspective consistent is essential!
  - the system & the surroundings
  - types of systems
    - open, closed, isolated
- Changes, Values & Size
  - What is change? (over time)
  - Measuring changes
    - State Functions vs Path Functions
  - Changes in Energy
    - Internal Energy
    - Heat & Work



- Energy Transfer/Change
  - by Heat
    - Heat (Temperature Change)
    - Heat Capacity
      - Specific Heat Capacity
      - Molar Heat Capacity
    - Heat Transfer
  - by Work
    - Pressure-Volume Work
  - Measuring Changes in Internal Energy
    - Bomb Calorimetry





Questions?