Slide 1

Thermochemistry

Energy, heat:
Uses and implications

Slide 2

Energy

Energy is actually like pornography: easy to recognize but hard to define.

Energy is sometimes defined as the capacity to do **work.**

Work is the result of a **force** acting over a distance.

Energy, however, takes many forms.

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Some Types of Energy

1. Light
2. Gravitational
3. Potential
4. Kinetic/mechanical
5. Chemical
6. Electrical
7. Magnetic
8. Heat – waste energy

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It's not what you say, it's what you do…

Often easier to see the effects of energy than the energy itself…

One of the easiest things to understand (especially during a Rochester winter) is "heat".

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Heat

Heat is actually waste energy from an engineering standpoint. It is energy that doesn't go into making the car move, or bonding the atoms together.

Heat simply raises the temperature of materials.

On the molecular level, it is energy that makes the molecules move around faster.

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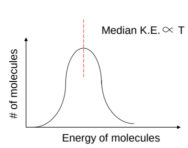
Kinetic Theory of Temperature

Within this model, "temperature" is actually a measure of the median kinetic energy of the molecules.



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When my water evaporated…

A. The temperature of the water stayed the same
B. The temperature of water stayed the same within the "uncertainty" of the measurement.
C. The temperature of the water went up.
D. The temperature of the water went down.
E. I have no frigging idea, the water fairies are looking better by the second.

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So…after my temperature "went down"…did

A. The evaporation stop?
B. The evaporation slow down?
C. The evaporation keep on trucking…?
D. Water fairies
E. Earth fairies

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Less water on my nightstand..

Is this because…

A. Water fairies stole it
B. The cat got it
C. The temperature in the room was briefly 100 degrees C
D. Some of the water molecules had enough energy to become gas molecules
E. I have no idea, I'm still asleep.

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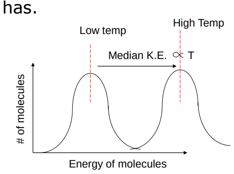
Absolute vs. Relative Temperature

Kinetic Energy is energy of motion. The faster a given object travels, the more KE it has.



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Absolute vs. Relative Temperature

Fahrenheit, Celsius, Kelvin

All different temperature scales.

You could define your own. Find two temperatures (body temperature, melting point of sugar), divide up the difference between them into arbitrary units and you're done!

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Absolute vs. Relative Temperature

Fahrenheit, Celsius were made just that way – picking two arbitrary temperatures and dividing the difference between them into arbitrary units.

These are "relative" temperature scales.

Relative scales work fine: higher temperature is "hotter", lower temperature is "colder"

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Absolute vs. Relative Temperature

But if Temperature is really to be defined as the KE of the molecules, 0 degrees should be the temperature at which ALL MOLECULES STOP MOVING!

A temperature scale with the correct “0” is called an “absolute” temperature scale. Kelvin is an absolute temperature scale!

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Energy is measured in Joules (J) which is a derived unit:

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

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How does a hot pack work?

When you break the seal, you mix water and a salt.

The process of dissolving the salt results in the release of heat.

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How does a cold pack work?

When you break the seal, you mix water and a salt together.

The process of dissolving the salt requires heat from the surroundings. Absorbing the heat cools the surroundings.

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Molecular Dynamics

How different are these?

$NaCl_{(s)}$

$NaCl_{(aq)}$

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Molecular Dynamics

How different are these?

$NaCl_{(s)}$ – solid salt

$NaCl_{(aq)}$ - salt dissolved in water

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Molecular Dynamics

Sometimes, they are doing something very dynamic but we still tend to think of them as static.

$NaCl_{(s)} \longrightarrow NaCl_{(aq)}$

When you see that reaction written, do you see it as a simple statement or a dynamic process?

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Molecular Dynamics

$NaCl_{(s)} \rightarrow NaCl_{(aq)}$

It's actually a very dynamic process. All of the Na^+ and Cl^- ions are ripped apart and surrounded by water molecules. There is energy associated with any such process: energy of attraction of ions, energy of attraction for the water and ions, etc.

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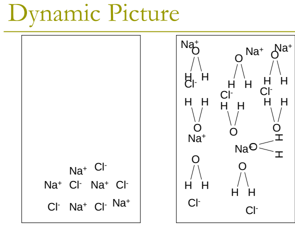
Dynamic Picture



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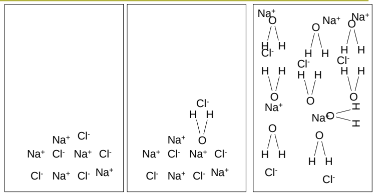
Dynamic Picture



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When I stick the waters in there…

A. Do I have to put more energy in
B. Do I get energy out
C. Is there no energy implications whatsoever

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Why do I get energy out?

A. I don't know, you just do…
B. My high school chem teacher told me so…
C. The electrostatic attraction is pleasing
D. There is a natural H to Na repulsion
E. There is a natural H to Na attraction

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What is step 2?

A. The Na and Cl get moved far enough apart to allow a water in
B. The water was ice and is melted
C. The water was liquid and was made a gas
D. I have no clue, stop wasting time with effing clicker questions
E. I really need to go the bathroom but I'm afraid to leave my clicker behind.

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I want to rip out a chloride ion…

A. I need to put energy in
B. I get energy out
C. I don't need to do anything with energy
D. Still asleep
E. Your mother.

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Dynamic Picture

$NaCl_{(s)} \rightarrow NaCl_{(aq)}$

To actually go from solid to aqueous requires individually separating each ion and then connecting each ion to the solvent network. It is the sum of billions and billions of little separations (a MOLE of separations).

There is actually a change in energy involved: the energy required to separate those billions and billions of molecules.

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$H_{2\,(g)} + O_{2\,(g)} \rightarrow H_2O_{(g)}$

How does this reaction occur? What do you need to do on the molecular level to make this reaction occur.

You need to break the H-H bond and the O=O bond, and make two O-H bonds!

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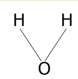
$H_{2\,(g)} + O_{2\,(g)} \rightarrow H_2O_{(g)}$

H-H

O=O

→



There is also an energy change involved in the reaction.

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Making and Breaking Bonds

Making a bond always releases energy.

Breaking a bond always requires energy.

Whether the entire reaction requires energy or releases energy depends on whether you get more/less energy out of the bonds you make than you put into the bonds you break.

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The System vs. The Surroundings

Energy accounting requires certain rules to maintain consistency and get appropriate units.

The System – what you are studying.

The Surroundings – the rest of the universe.

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Kinds of systems

There are different kinds of systems, defined by how they are related to the universe:

Open system – directly connected to the universe, mass and heat can go back and forth between system/universe.

Closed system – no mass can be transferred, but heat can be.

Isolated system – no mass or heat can be transferred.

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Types of System

An "open system" would be this room.

A "closed system" would be an unopened bottle of wine.

An "isolated system" would be a sealed perfect thermos – your soup stays hot for an eternity!

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Approximations to the Ideal

Ideally, we would always have an isolated system to study. It is much simpler if everything we are studying is trapped.

In practice, we only have nearly isolated systems.

For example, a thermos. (Sorry, but it won't keep your soup warm for an eternity.)

As long as it limits heat loss or the duration of the experiment, it is practically isolated.

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Calorimetry

Determining "heat" by measuring temperature.

Adding heat to something raises its temperature. That temperature rise is quantitatively related to the amount of heat added.

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Boiling water

Is it easier to boil a cup of water or boil a swimming pool?

The bigger amount (mass) of water requires more energy to get the same temperature change!

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Is it easier to boil a cup of water or boil a swimming pool

A. Cup of water
B. Swimming Pool
C. They're the same
D. I wouldn't do either one.
E. Why are we clicking on this?

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Specific Heat

Have you ever put a ceramic pan on the stove and a metal pan on the same stove? Is there a difference?

The ceramic pan usually heats up more slowly (and cools down more slowly) than the metal pan.

Adding the same amount of heat to different materials causes a different temperature change!

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Specific Heat

The amount of heat added to a fixed amount (mass) of a substance to change the temperature 1° is called the "specific heat of the substance"

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The relation of Heat to Temperature

Putting it all together:

$$q = m\ c\ \Delta T$$

q = heat
m = mass of the object
c = specific heat of the object
Δ T = change in temperature

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Calorimetry problem

I want to heat 1 gallon of water from 20 °C to 30°C. How much heat must I add?

$$q = m\ c\ \Delta T$$

$1 \text{ gal} * \frac{3.7854 \text{ L}}{1 \text{ gal}} * \frac{1000 \text{ mL}}{1 \text{ L}} * \frac{1 \text{ g}}{1 \text{mL}} = 3785.4 \text{ g}$

c = 4.18 J/g °C FOR WATER
$\Delta T = T_f - T_i = 30\ °C - 20\ °C = 10\ °C$

q = (3785.4 g) (4.18 J/g °C) (10 °C) = +158, 230 J
= 158.230 kJ

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Calorimetry Problem #2

I have 50.0 mL of pure water in a perfectly insulated thermos that is at room temperature (25 ° C). I put a 5.0 g aluminum slug into a dry test tube and then put it in boiling water at sea level for several minutes until the aluminum slug, test tube and water have all reached equilibrium. I then instantaneously dump the hot slug into the perfectly insulated thermos without splashing. What is the temperature of the water and slug when they reach equilibrium with each other BUT NOT THE ROOM?

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Solution to Question #2

$q_{water} = - q_{Al}$
$q_{water} + q_{Al} = 0$

$q_{water} = m_{water}\, c_{water}\, \Delta T_{water}$
$q_{Al} = m_{Al}\, c_{Al}\, \Delta T_{Al}$

$m_{water}\, c_{water}\, \Delta T_{water} = - m_{Al}\, c_{Al}\, \Delta T_{Al}$

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Solution to Question #2

50.0 mL water * $\frac{\underline{1.0 \text{ g water}}}{1.0 \text{ mL water}}$ = 50.0 g water

From the table in the book:

c_{water} = 4.18 J/g ºC
c_{Al} = 0.90 J/g ºC

$m_{water}\, c_{water}\, \Delta T_{water} = - m_{Al}\, c_{Al}\, \Delta T_{Al}$
50.0 g * 4.18 J/g ºC ($T_{f,water}$-25) =
- 5.0 g * 0.90 J/g ºC * ($T_{f,Al}$ – 100)

$T_{f,water} = T_{f,Al}$

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Solution to Question #2

50.0 g * 4.18 J/g ºC (T_f-25) =
- 5.0 g * 0.90 J/g ºC * (T_f – 100)

209 (T_f-25) = -4.5(T_f – 100)
209 T_f – 5225 = -4.5T_f + 450
209 T_f + 4.5 T_f = 450 + 5225
213 T_f = 5675
T_f = 26.58 ºC

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C=4.18 J/gC (water)
c=0.90 J/gC (Al)

I take 50 mL of water at 25ºC and add to it a test tube of 25.0 mL of boiling water which has a 5.0 g Al slug in it. What is the final Temperature?

A. 27
B. 57
C. 24
D. 161
E. 42.54

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Question

Burning 1 g of gasoline releases 3100 Joules of energy. How much gasoline (grams) would I need to burn to boil 1 L of water that is initially at room temperature (298 K)?

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$$\frac{3100\,J}{g\ burned}$$

Burning 1 g of gasoline releases 3100 Joules of energy

UNITS! UNITS! UNITS

$$\frac{3100\,J}{g\ burned}$$

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$$\frac{3100\ J}{g\ burned}$$

How much gasoline (grams) would I need to burn to boil 1 L of water that is initially at room temperature (298 K)?

How much heat does it take?

If I'm changing the T of water, it's:

$Q=mc\Delta T$

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$$\frac{3100\ J}{g\ burned}$$

How much gasoline (grams) would I need to burn to boil 1 L of water that is initially at room temperature (298 K)?

$$1\ L\ \frac{1000\ mL}{L}\ \frac{1\ g}{1\ mL} = 1000\ g\ water$$

$$Q = (1000\ g)\left(4.18\frac{J}{gK}\right)(373K - 298K)$$

$$= 313{,}500\ J$$

$$313{,}500\ J\ \frac{1\ g\ burned}{3100\ J} = 101\ g\ gas\ burned$$

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Another little question

I burn 100 g of ethanol (C_2H_5OH) in a bomb calorimeter which contains 1 L of water at 298.0 K. After combustion of the ethanol, the temperature of the water in the bomb has climbed to 304.6 K. What is the energy released per mol of ethanol (assume negligible heat absorption by the bomb)?

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What’s a bomb calorimeter?

Dang…I need the blackboard – I’ll draw this one out tomorrow.

But the bomb is a bomb-proof container that holds a combustion reaction. The container is submerged in a water bath which acts as a heat reservoir, kind of like this…

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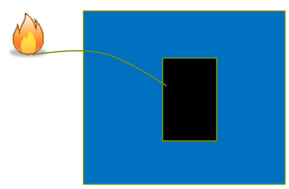

Slide 54

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If T changed…then…

$$Q = mc\Delta T = (1000\ g)(4.18 \frac{J}{gK})(304.6K - 298K)$$

$$Q = 27{,}588\ J water = -27{,}588\ J\ ethanol$$
$$= -27.588\ kJ$$

$$100\ g\ C_2H_5OH \frac{1\ mol}{46.07\ g} = 2.2\ mol\ C_2H_5OH$$

$$\frac{-27.588\ kJ}{2.2\ mol} = -12.54 \frac{kJ}{mol} = \Delta H_{comb}$$