Ch03

Energy

The capacity to do work.

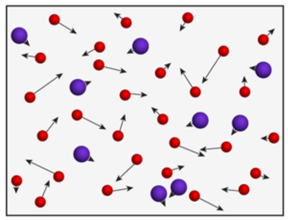
How matter moves.

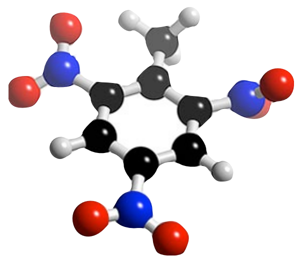
version 1.5

Ch03

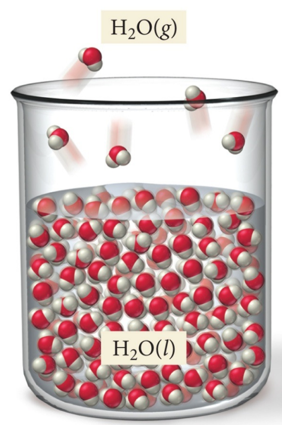
Energy Affecting Matter

- Energy
  - Defined
  - Kinetic vs Potential
    - Thermal vs Chemical
  - Units of Energy
- Energy Content
  - Energy value
  - Total energy
- Temperature
  - Temperature scales
    - Celsius, Fahrenheit, Kelvin
    - Converting
  - Measurement Types
    - Size, Position, Change
    - Finding ΔX



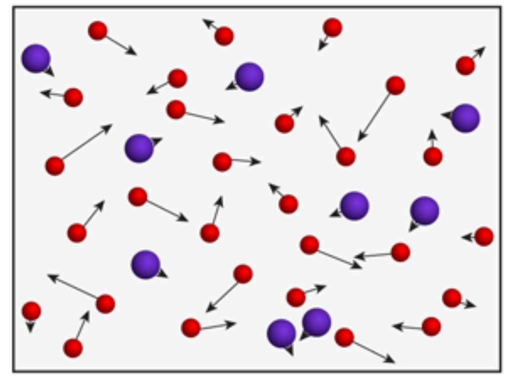


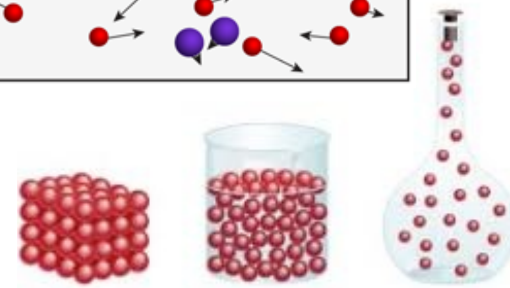
- Energy Affecting Matter
  - Heat
    - Heat Capacity
    - Example Calculations
  - State
    - Critical Temperatures
    - Example Calculations

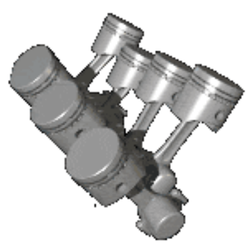


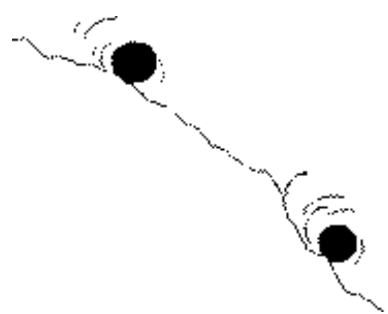
Energy

- Energy (E) is the capacity to do work or transmit heat.
  - Work is a force moving an object (matter).
  - Energy is what makes things move.



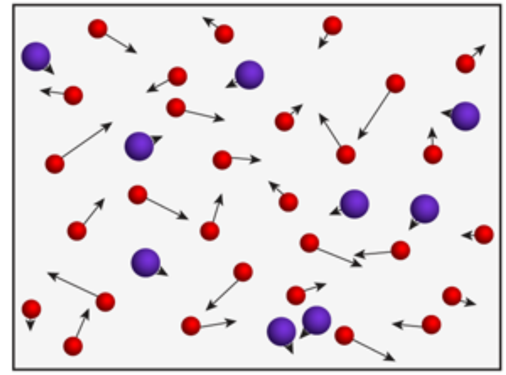


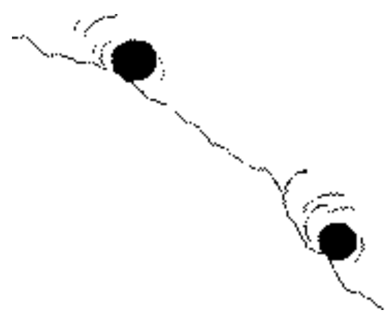




Energy

- 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.





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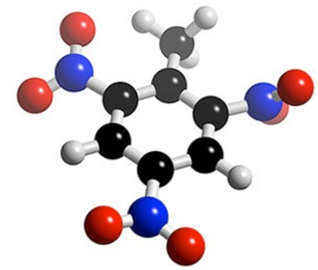

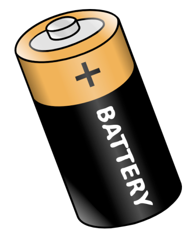

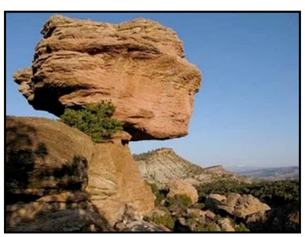

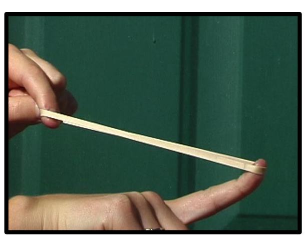

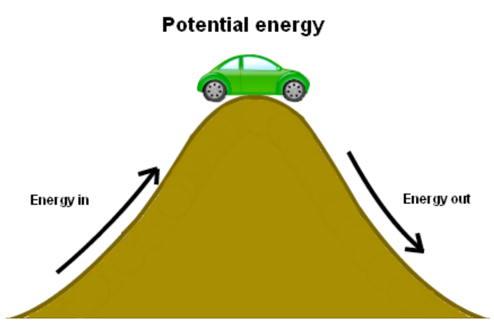

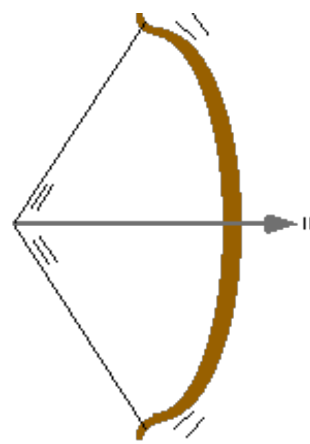

Units of Energy

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.

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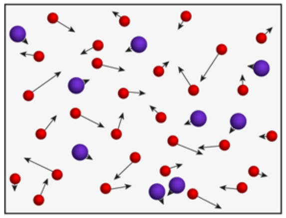
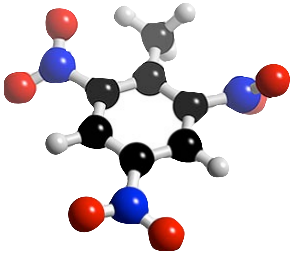
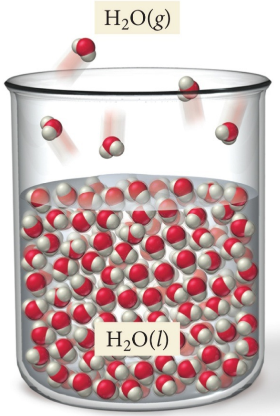
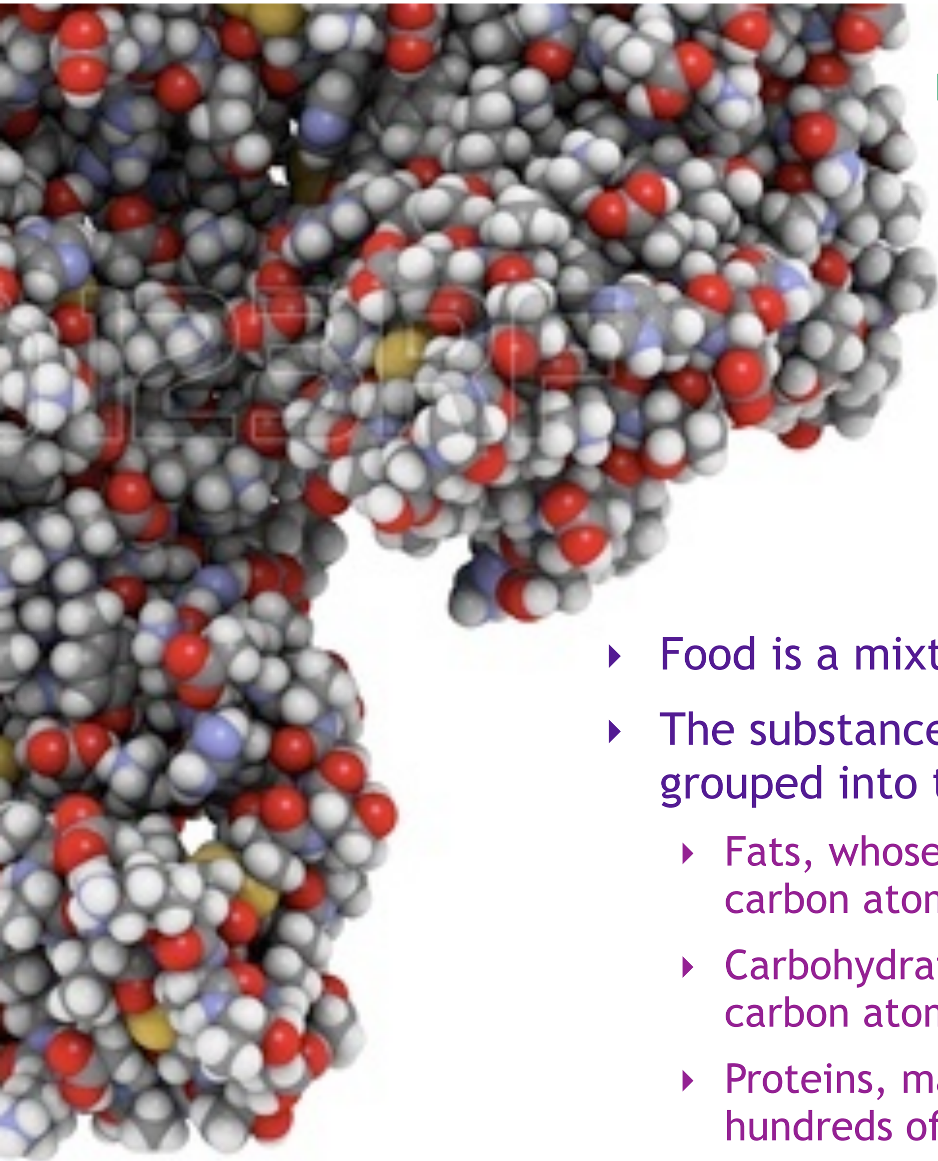



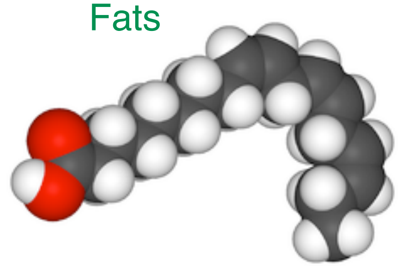





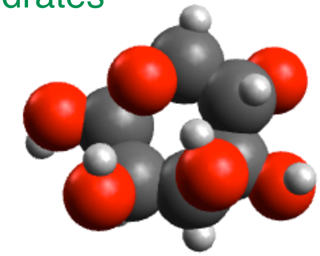


Proteins

Fats



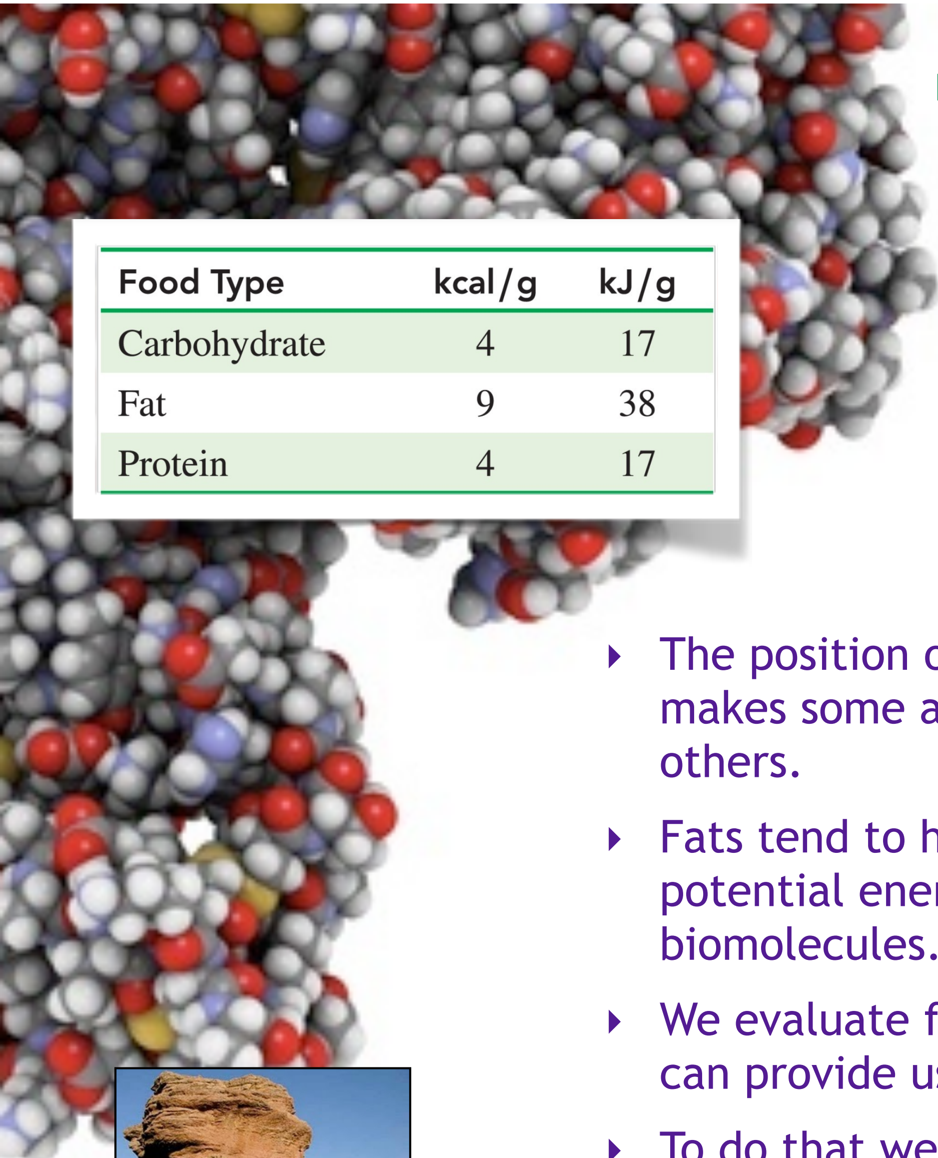
Carbohydrates



- Food is a mixture.
- The substances that make up most food can be grouped into three types of matter:
  - Fats, whose particles are chains of about 12-20 carbon atoms.
  - Carbohydrates, particles that are rings of 5 or 6 carbon atoms with lots of oxygen atoms attached.
  - Proteins, massive particles that are often made up of hundreds of thousands of atoms each.

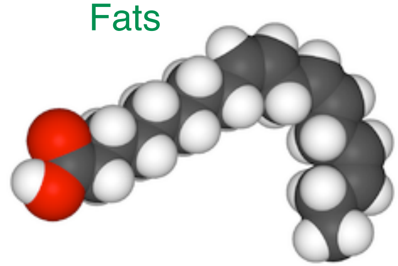
- Each of these types of molecules serve different purposes in the bodies biochemistry.
- We’ll get into more details about each of these kind of biomolecule next semester.

Proteins

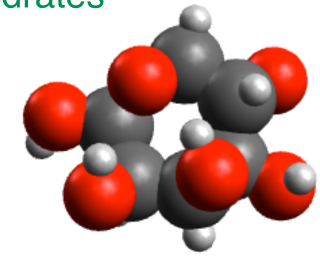


Food Type	kcal/g	kJ/g
Carbohydrate	4	17
Fat	9	38
Protein	4	17

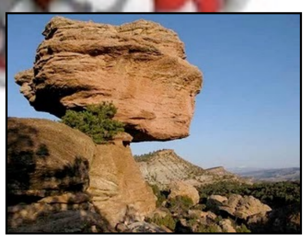
Fats

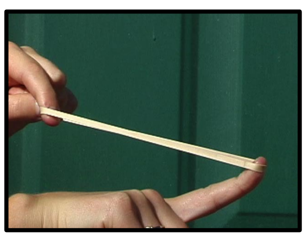


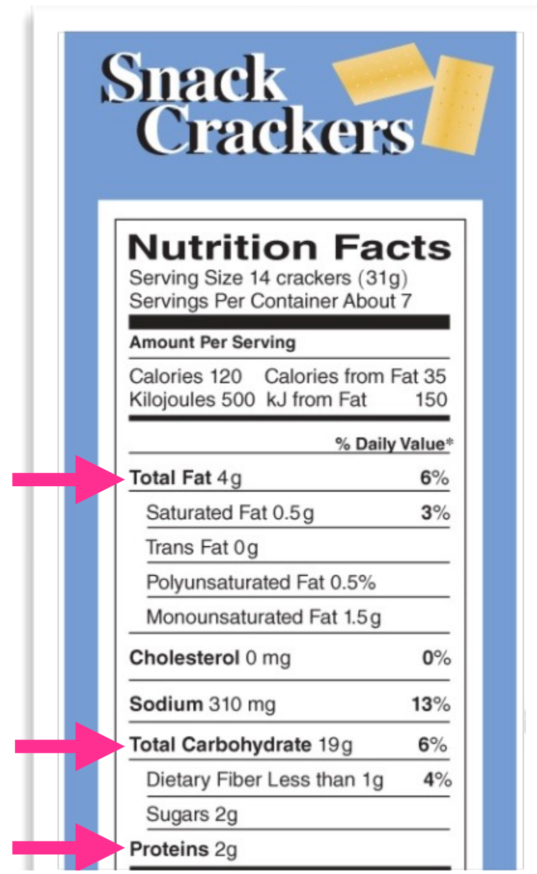
Carbohydrates



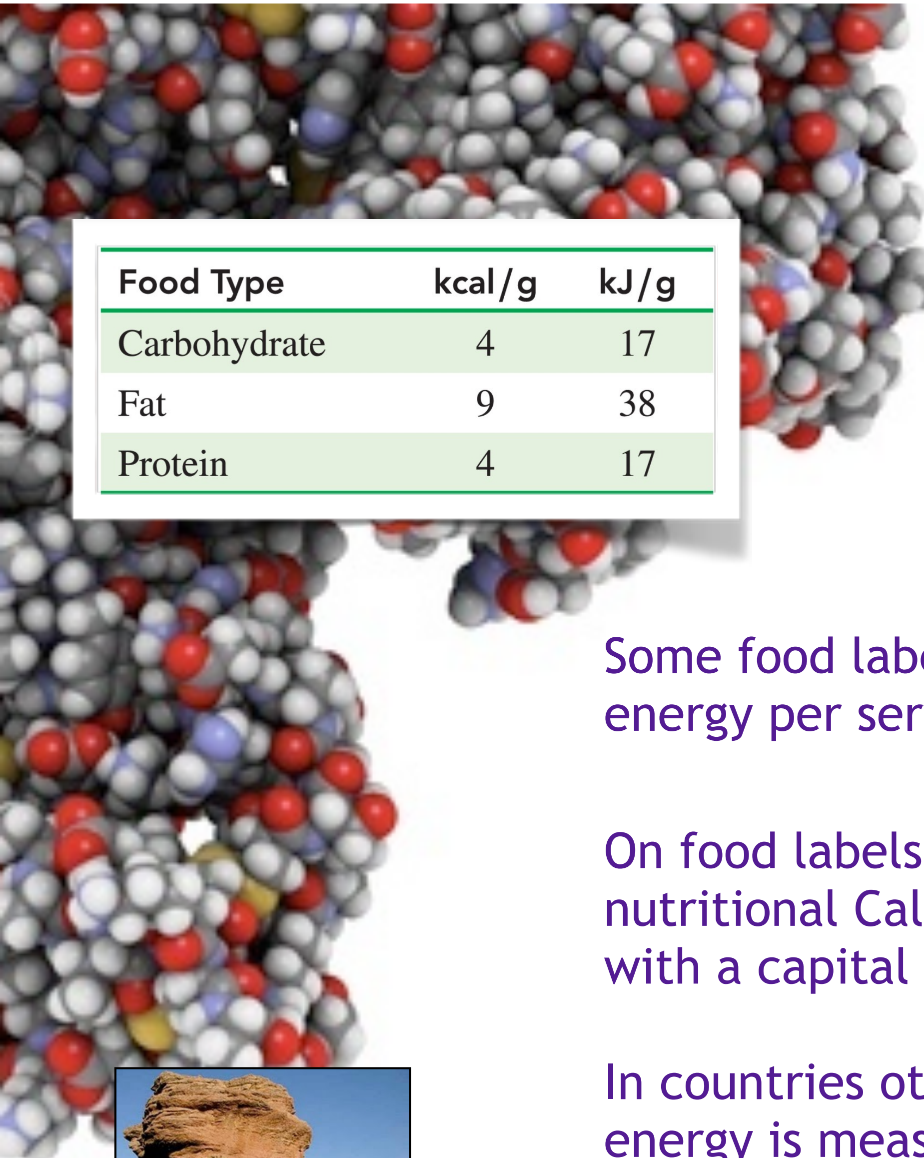
- The position of atoms in these particles makes some a better source of energy than others.
- Fats tend to have more than twice the potential energy per gram of the other biomolecules.
- We evaluate food by how much energy it can provide us.
- To do that we look at how much matter in each of these three categories it contains.







Proteins



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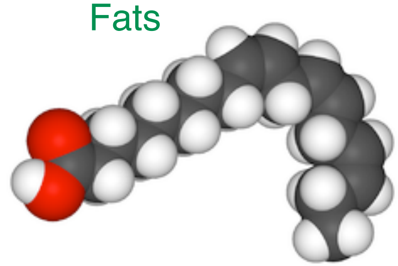

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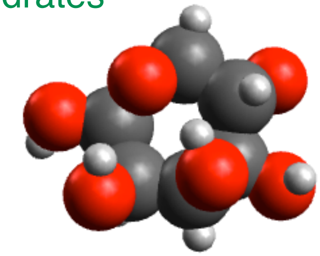
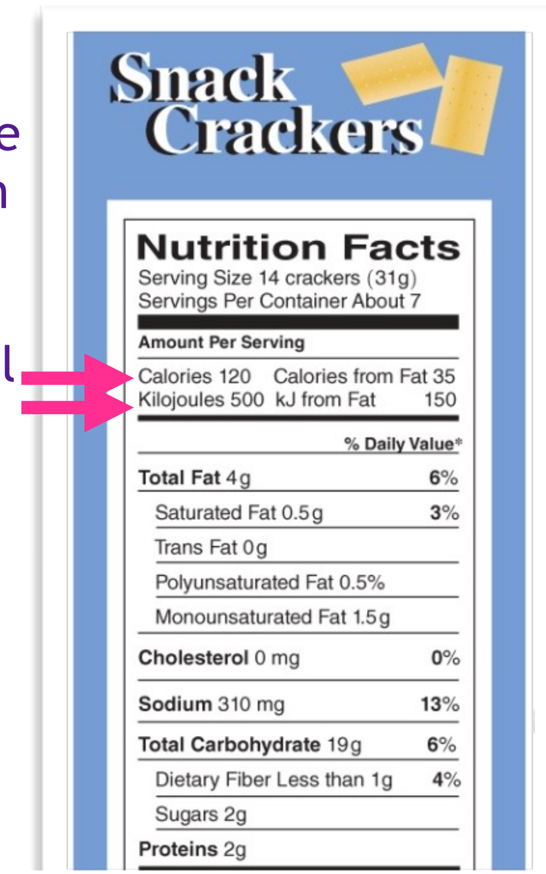
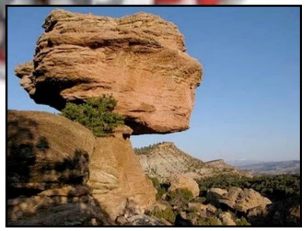


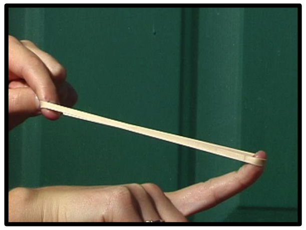
Some food labels will report the total energy per serving of that substance.

On food labels, energy is reported with the nutritional Calorie (or big calorie), written with a capital C.

In countries other than the U.S. nutritional energy is measured in kilojoules (kJ).







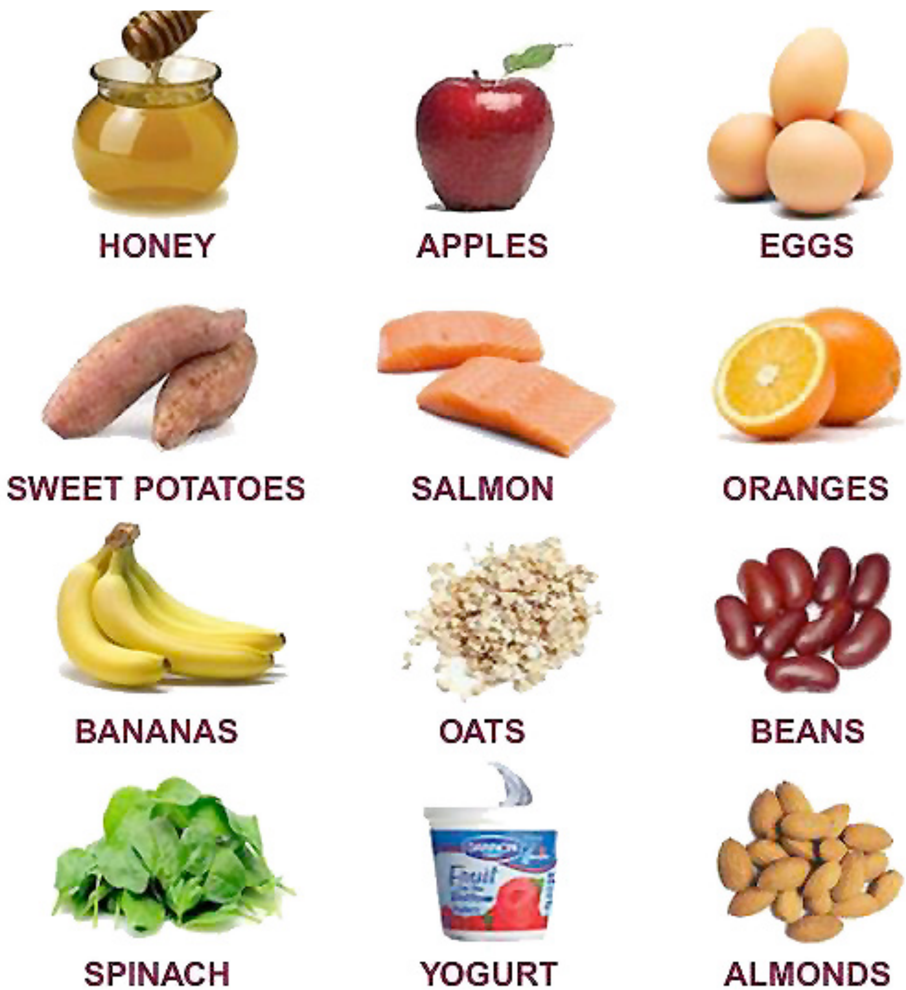
1 Cal = 1000 calories

1 Cal = 1 kcal

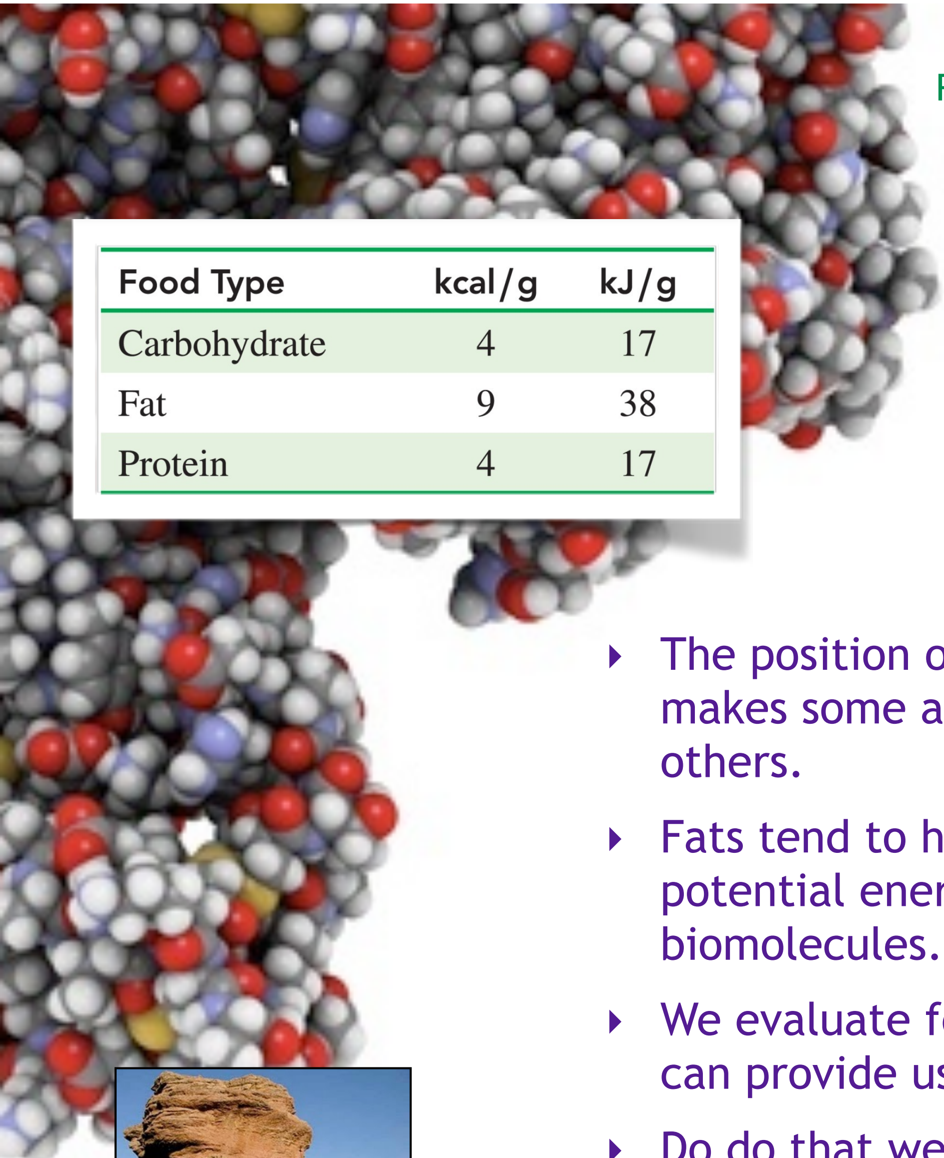
1 kJ = 1000 J

Energy in Food

- The energy value (caloric value) for 1 g of a food is given in kilojoules (kJ) or kilocalories (kcal).



Food	Carbohydrate (g)	Fat (g)	Protein (g)	Energy*
Banana, 1 medium	26	0	1	110 kcal (460 kJ)
Beef, ground, 3 oz	0	14	22	220 kcal (910 kJ)
Carrots, raw, 1 cup	11	0	1	50 kcal (200 kJ)
Chicken, no skin, 3 oz	0	3	20	110 kcal (460 kJ)
Egg, 1 large	0	6	6	80 kcal (330 kJ)
Milk, 4% fat, 1 cup	12	9	9	170 kcal (700 kJ)
Milk, nonfat, 1 cup	12	0	9	90 kcal (360 kJ)
Potato, baked	23	0	3	100 kcal (440 kJ)
Salmon, 3 oz	0	5	16	110 kcal (460 kJ)
Steak, 3 oz	0	27	19	320 kcal (1350 kJ)



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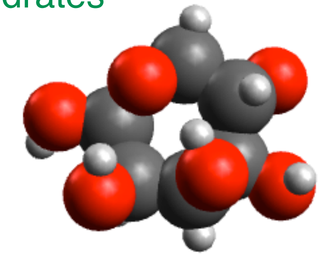
Proteins

Carbohydrates



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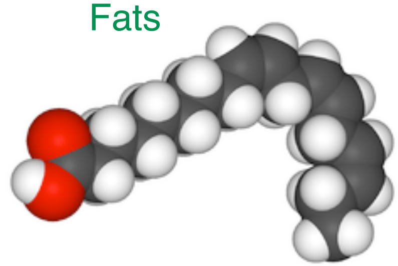

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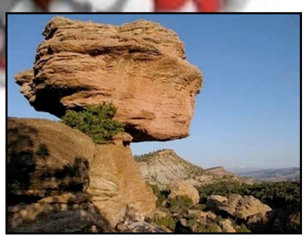

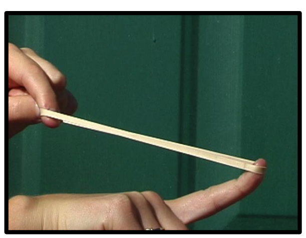

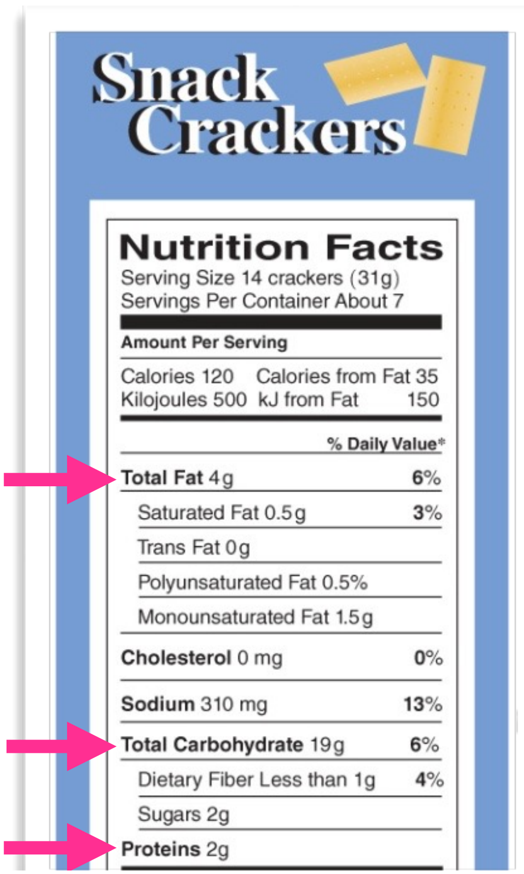

Energy content of a cup of milk.

A cup of whole milk contains 13 g of carbohydrate, 9.0 g of fat, and 9.0 g of protein. What is the energy content of a cup of milk (in kilocalories)?

Carbohydrates have an average energy value of 4.0 kcal/g, fat is 9.0 kcal/g, and protein is 4.0 kcal/g.

Carbs $13 g \cdot 4.0 \frac{kcal}{g} = 52 \text{ kcal}$

Fat $9.0 g \cdot 9.0 \frac{kcal}{g} = 81 \text{ kcal}$

Prot $9.0 g \cdot 4.0 \frac{kcal}{g} = 36 \text{ kcal}$

+

Total 169

169 kcal

Health Facts

The number of kilocalories or kilojoules needed in the daily diet of an adult depends on gender, age, and level of physical activity.

A person loses weight when food intake is less than energy output.

Energy Requirements for an Adult

Gender	Age	Moderately Active kcal (kJ)	Highly Active kcal (kJ)
Female	19–30	2100 (8800)	2400 (10 000)
	31–50	2000 (8400)	2200 (9200)
Male	19–30	2700 (11 300)	3000 (12 600)
	31–50	2500 (10 500)	2900 (12 100)

How we use Energy

Activity	Energy (kcal/h)	Energy (kJ/h)
Sleeping	60	250
Sitting	100	420
Walking	200	840
Swimming	500	2100
Running	750	3100

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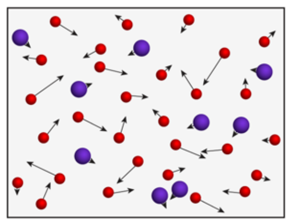

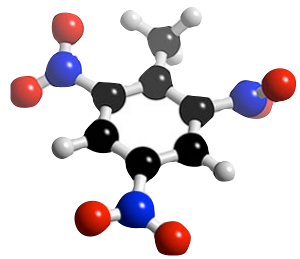

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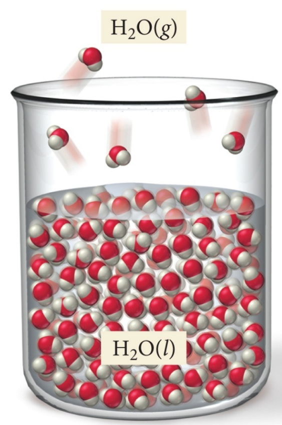
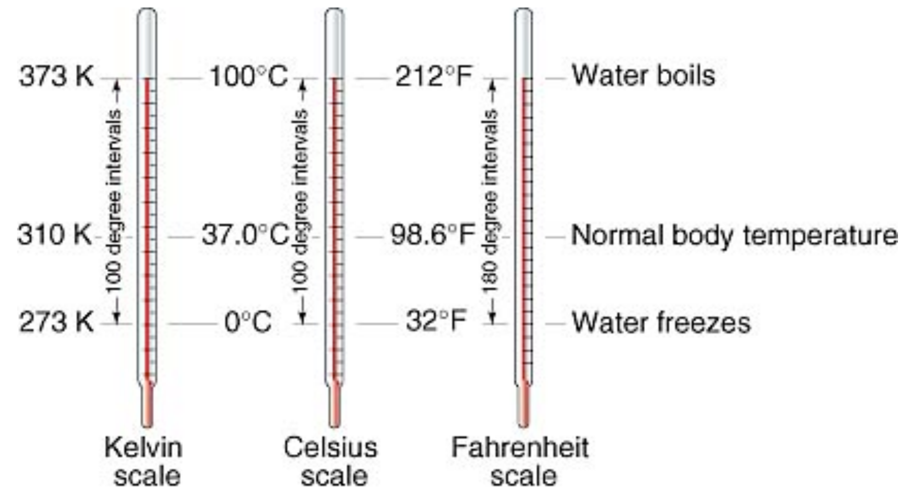
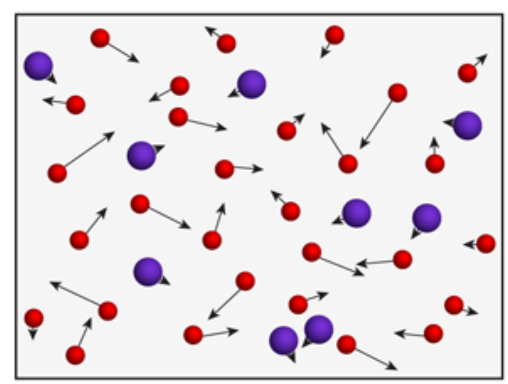


Temperature

- Temperature is an intensive property.
- Temperature is a measure of the average kinetic energy of the particles in a substance (the average thermal energy).
  - It tells you, on average, how much thermal energy it's particles have.
- Temperature is measured on three different scales, with difference reference points.
  - on the Celsius scale
    - The reference point is the temperature at which water freezes, it's defined as 0˚C.
  - on the Fahrenheit scale
    - The reference point is the temperature at which brine (salt) water freezes, it's defined as 0˚ F.
    - Fahrenheit thermometers are now more often calibrated by the freezing point of water 32˚ F.
  - on the Kelvin scale
    - The reference point is the projected temperature at which an ideal gas would have zero volume – 0 K.
    - We'll discuss this more when we get to the gas state of matter (chapter 8).





Temperature

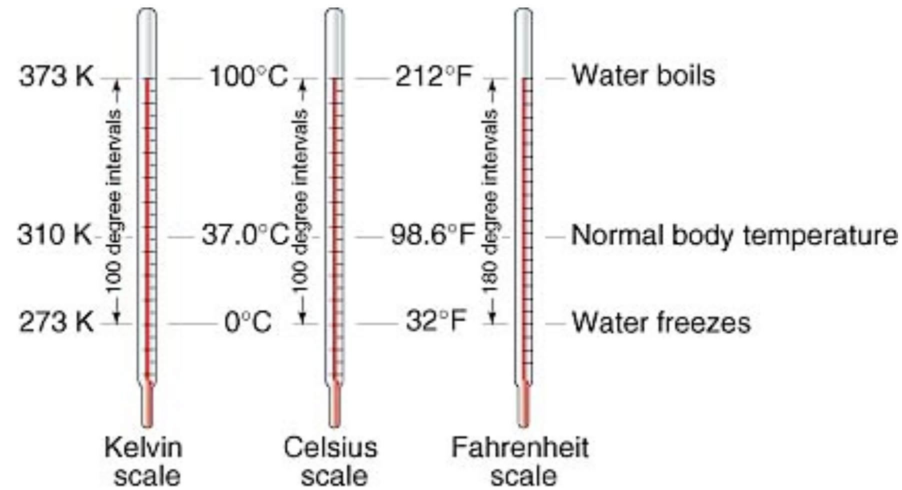
- The size of a Celsius and Kelvin unit of temperature is the same.
  - So a change of 100˚C is equal in size to a change of 100 K.
  - Converting between Celsius and Kelvin scales is just changing the reference point by 273.15.

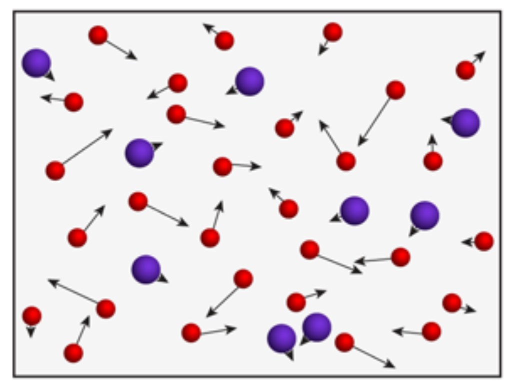
$$K = C + 273.15$$

- The size of a Fahrenheit unit of temperature not (a Fahrenheit degree is 1.8 times larger).
  - So a change of 180˚C is equal in size to a change of 100 ˚F.
  - Converting to Fahrenheit is more complex.
    - You need to multiply the celsius measure by 1.8
    - You need to change the reference point by 32 degrees

$$F = C\,(1.8) + 32$$

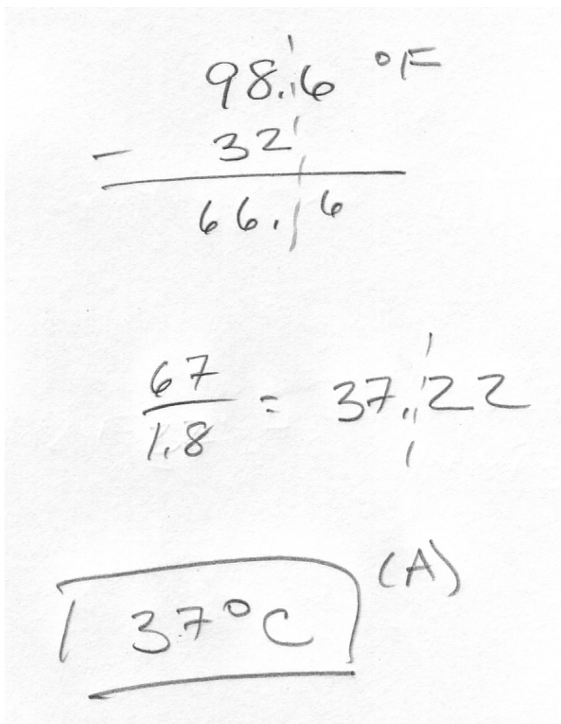
$$C = (F - 32)/1.8$$

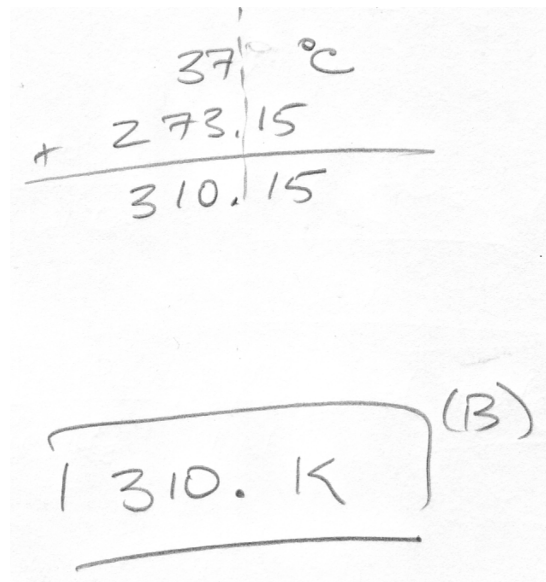




Body Temperature

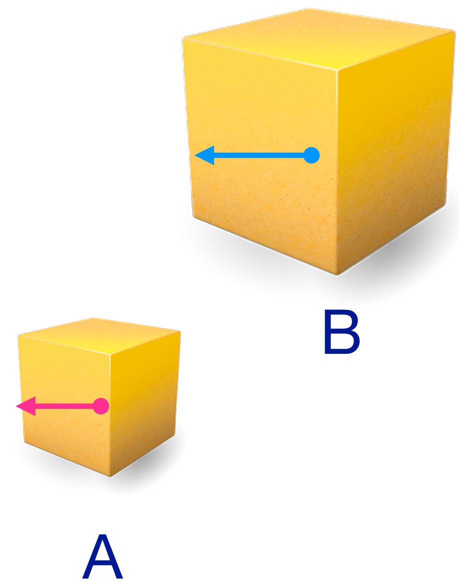
Normal body temperature is 98.6 ˚F. (a) What is this measure in Celsius? (b) What does this temperature measure in Kelvin?

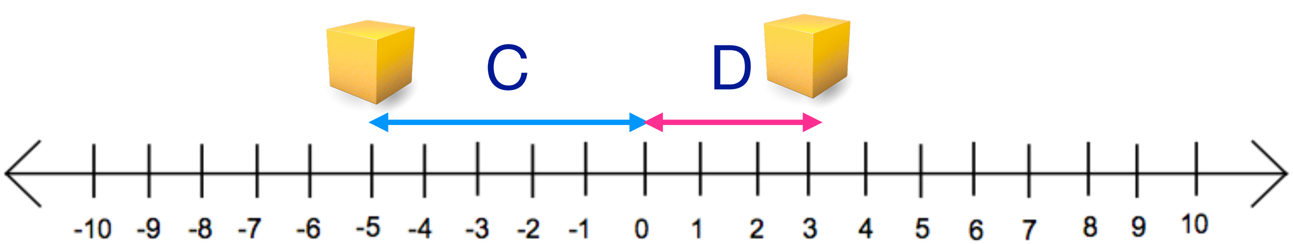




Measurements of Size & Position

- 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 move forward.
  - So change is always final minus initial.
- The change of a value is represented with the delta symbol (Δ).

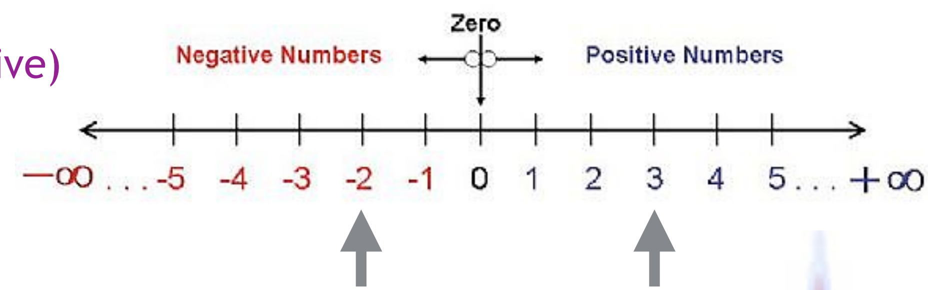
X_{FINAL} is -2.

$X_{INITIAL}$ is +3.

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

- 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

$\Delta X = -5$

Change in Temperature

A block of metal cools from 47.1 ˚C to 26 ˚C what is the change in temperature?
(A) In Celsius? (B) In Kelvin?

$\Delta T = T_F - T_I$

$= 26.°C - 47.1°C$

$= 21°C$

$\Delta T = -21°C$ (A)

26 °C
$- 47.1$
-21.1 °C

$\Delta T = T_F - T_I$

$= 299 K - 320.3 K$

$= -21 K$

$\Delta T = -21 K$ (B)

T_F 26 °C
$+ 273.15$
299.15 K

T_I 47.1 °C
$+ 273.15$
320.25 K

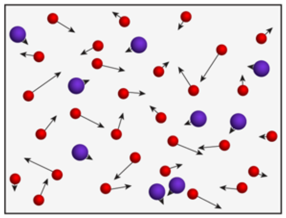
ΔT 299 K
$- 320.3$ K
21 K

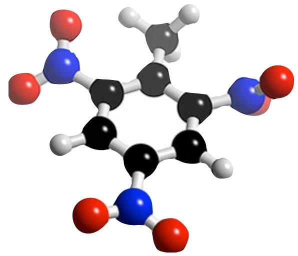
Ch03

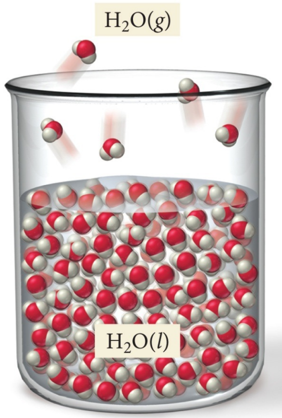
Energy Affecting Matter

- Energy
  - Defined
  - Kinetic vs Potential
    - Thermal vs Chemical
  - Units of Energy
- Energy Content
  - Energy value
  - Total energy
- Temperature
  - Temperature scales
    - Celsius, Fahrenheit, Kelvin
    - Converting
  - Measurement Types
    - Size, Position, Change
    - Finding ΔX

- Energy Affecting Matter
  - Heat
    - Heat Capacity
    - Example Calculations
  - State
    - Critical Temperatures
    - Example Calculations

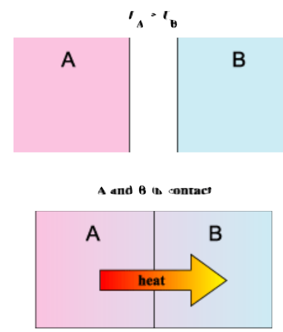


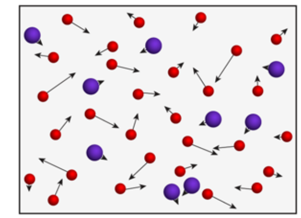




Heat

- Heat (q or Q) is energy that flows between two systems due to their difference in temperature.
- Heat is a change in energy (ΔE).
- Heat into an object increases the average energy of it's particles, it increases the temperature of the object.

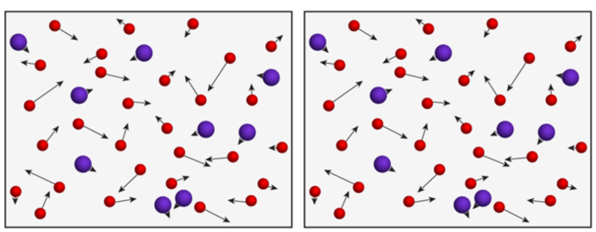


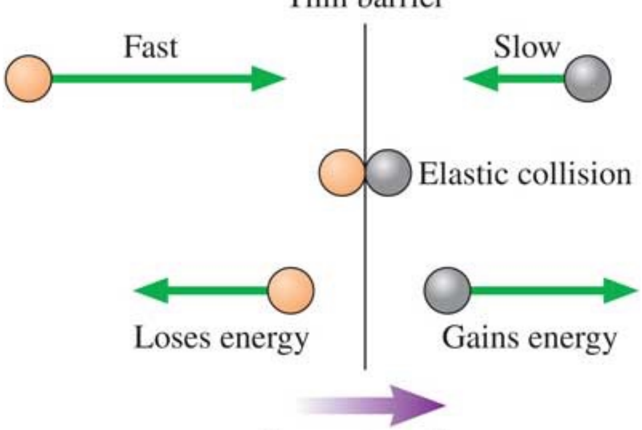


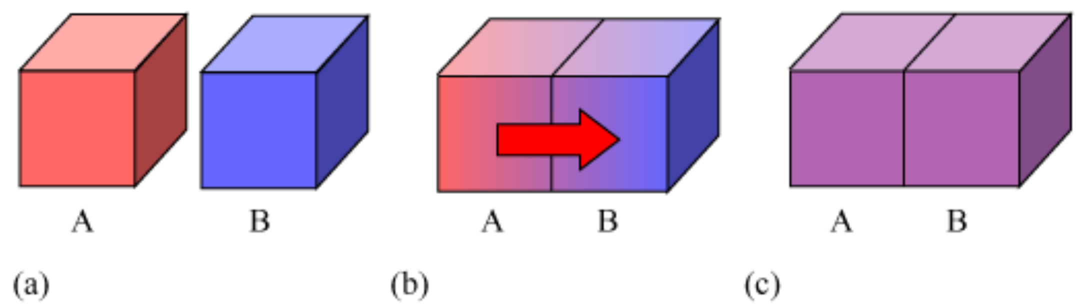
↑E ↑T

ΔE α ΔT

q α ΔT

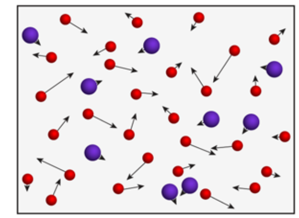


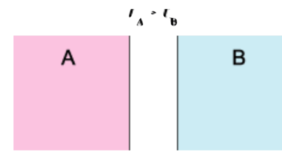


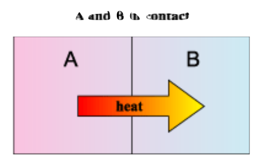


Heat is a Change

- Heat (q or Q) is energy that flows between two systems due to their difference in temperature.
  - Heat is a change in energy (ΔE).
- Heat into an object increases the average energy of it's particles, it increases the temperature of the object.
- A physical change occurs to A and B when energy is lost or gained.
- Changes that result from loosing energy are exothermic.
  - A undergoes an exothermic change (it releases energy).
- Changes that result from gaining energy are endothermic.
  - B undergoes an endothermic change (it gains energy).





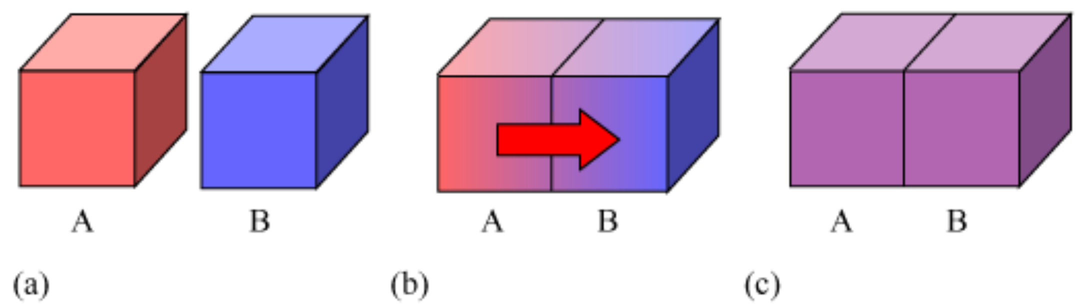


$\uparrow E \quad \uparrow T$

$\Delta E \quad \alpha \quad \Delta T$

$q \quad \alpha \quad \Delta T$

For block "A"
q is Exothermic



For block "B"
q is Endothermic

Endothermic vs Exothermic

Tips:

- Don't think about the **temperature (T)** (how hot or cold it is).
- Think about the **heat (q)** (thermal energy going in or out).
- Endothermic/Exothermic describes the direction of the heat.
- It describes action, what is being ***done***.

Ice Melting

Endothermic

Wood Burning

Exothermic

Breath Freezing

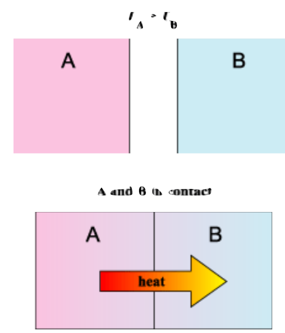
Exothermic

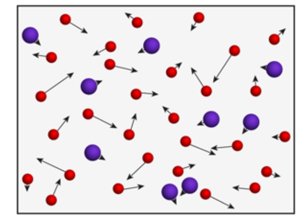
Water Boiling

Endothermic

Heat Capacity (of an Object)

- 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

ΔE α ΔT

q α ΔT

- Heat capacity (C) is the energy 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.

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

C = 2,835 J / ˚C

C = 124.0 J / ˚C

C = 62.0 J / ˚C

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

Specific Heat Capacity (of a Substance)

- **Heat capacity** (C) is the energy 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.

There are two forms of the heat capacity equation!

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

- This gives us two new kinds of heat capacity:
  - **Specific Heat Capacity** (C_s) is the heat capacity per gram (m) of a substance.
    (Often abbreviated as Specific Heat)

$$C_s = C \;/\; m \quad \text{or} \quad C = C_s * m$$

- 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.

C = 2,835 J / ˚C

C_s = 0.45 J / g ˚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

$\Delta T = T_F - T_I$

$= 28.7°C - 53.2°C$

$= -24.5\ °C$

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

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

$= -5818.75\ J$

$= -5.8\ kJ$

28.7
− 53.2
−24.5

The energy in the system went down by 5.8 kJ

+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

$$q = C_s \cdot m \cdot \Delta T$$
$$= 4.184\ J/g°K \cdot 143\ g \cdot 53.0\ K$$
$$= 3.17 \times 10^4\ J$$

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

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

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

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

$C_s = 0.450\ J/gK$

$m = 149g$

$T_I = 62.0°C$

$\Delta T = T_F - T_I$

$T_F = ?$

Calculate ΔT ; Solve for T_{FINAL}

Part A

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

$\Delta T = \frac{q}{C_s \cdot m}$

$= \frac{-2.34 \times 10^3\ J}{0.450\ J/gK \cdot 149g}$

$= -34.9K$

for a change in T

$\Delta T = -34.9°C$

Part B

$\Delta T = T_F - T_I$

$T_F = \Delta T + T_I$

$= -34.9°C + 62.0°C$

$= 27.1°C$

$27.1°C$

Ch03

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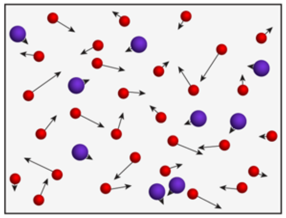
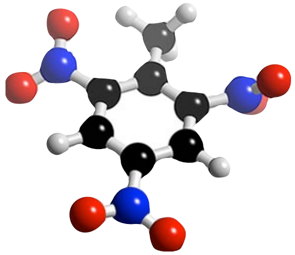
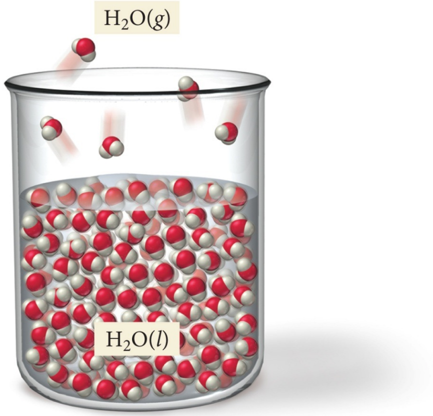
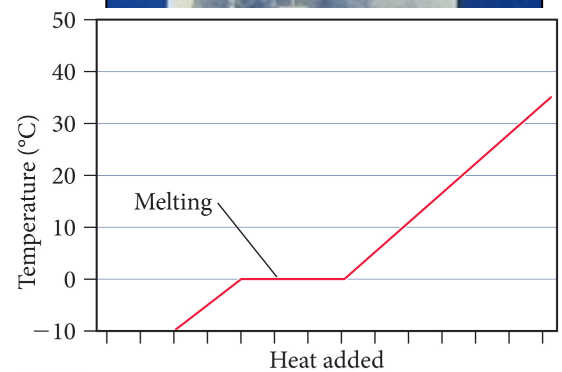


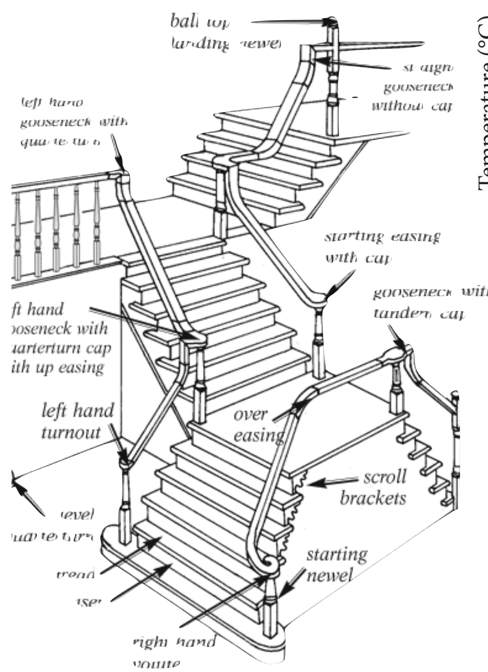




Melting

- Melting point is a second threshold.
- Melting point is where the intermolecular forces that lock the substance into a rigid solid are overcome by the heat added.
- It's when the substance has absorbed all the heat it can possibly hold in that rigid solid structure.
- Every joule of heat added after that, takes some portion of solid to a liquid — at the same temperature.
- Ice baths:
  - An ice bath contains water at 0˚C and ice at 0˚C.
    - Add more heat, more of it becomes water — but the temperature doesn't change.
    - Remove heat, more of it becomes ice — but the temperature doesn't change.
  - That's why we use ice baths — to insure a constant, reliable temperature.

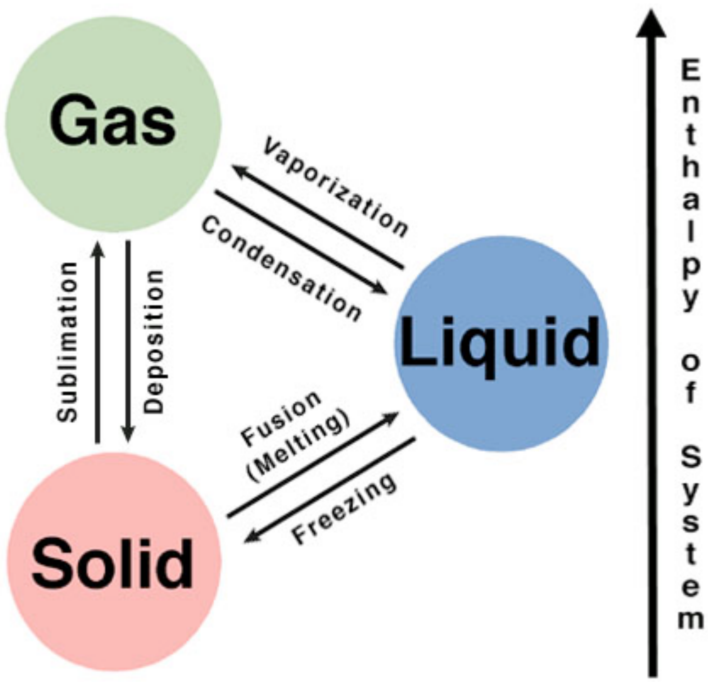




State Changes

- There are three states of matter and six processes by which we change state.
- Three are endothermic processes, three are complementary exothermic processes.

Note: The process of melting is described with Heat of Fusion. (It's an old convention)



Endothermic

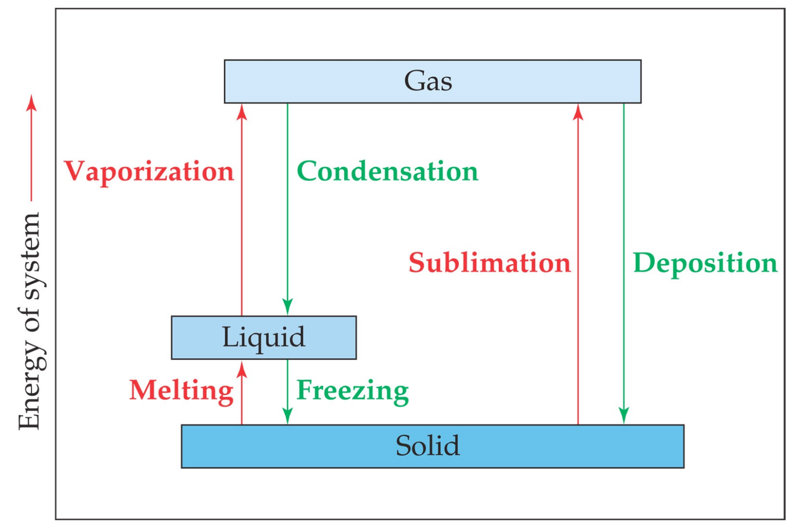
- Fusion
- Vaporization
- Sublimation
- Condensation
- Freezing
- Deposition

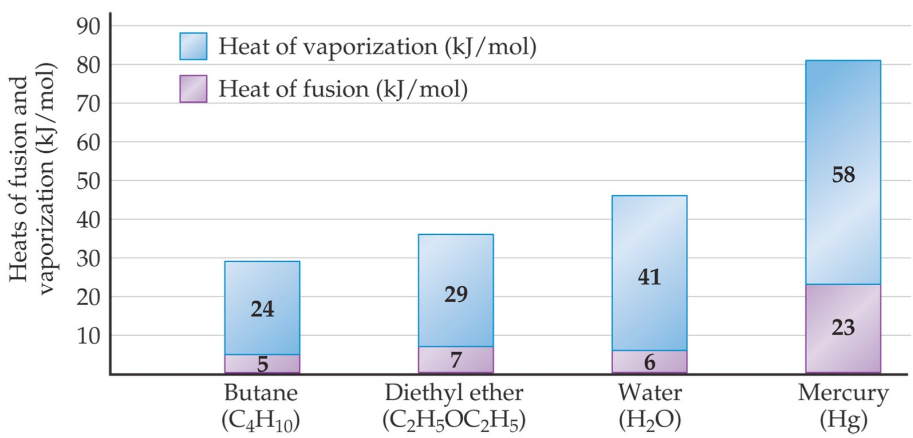
Exothermic

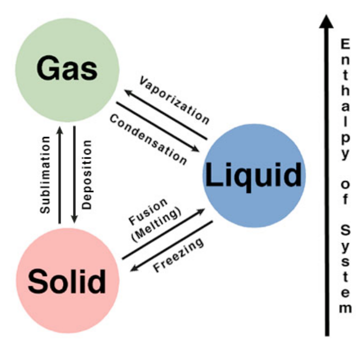
State Changes

- There are three states of matter and six processes by which we change state.
- Three are endothermic processes, three are <u>complementary</u> exothermic processes.
- Because of Hess's Law, Heat of Sublimation is equal to the sum of the heat of Vaporization and Fusion.

- Fusion(Melting) ΔH_{fus} 334 J/g
- Vaporization ΔH_{vap} 2,257 J/g
- Sublimation ΔH_{sub} 334 + 2,257 = 2,591 J/g
- Freezing ΔH_{frz} = $-\Delta H_{fus}$ = -334 J/g
- Condensation ΔH_{con} = $-\Delta H_{vap}$ = -2,257J/g
- Deposition ΔH_{dep} = $-\Delta H_{sub}$ = -2,591 J/g







Melting Ice

A beaker has 243.1 grams of water in it. The heat of fusion of water is 2,257 J/g. What is the change in energy (q) of that water when it freezes?

heat of fusion of water = 334 J/g
heat of vaporization of water = 2257 J/g
specific heat of ice = 2.09 J/g·°C
specific heat of water = 4.18 J/g·°C
specific heat of water vapor = 1.99 J/g·°C

$$\Delta H_{FRZ} = -\Delta H_{FUS} = -2{,}257 \text{ J/g}$$

$$q = m \cdot \Delta H_{FRZ}$$

$$= 243.1 \text{ g} \cdot (-2{,}257 \text{ J/g})$$

$$= -548{,}676.7 \text{ J}$$

$$= -5.487 \times 10^{5} \text{ J}$$

(4 s.f.)

Energy Affecting Matter

- Energy
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    - Thermal vs Chemical
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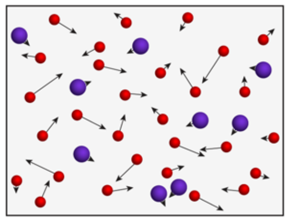
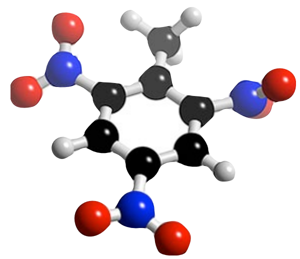
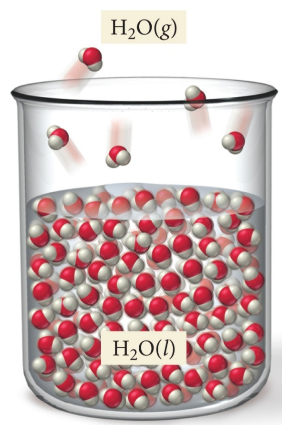




- Energy Affecting Matter
  - Heat
    - Heat Capacity
    - Example Calculations
  - State
    - Critical Temperatures
    - Example Calculations

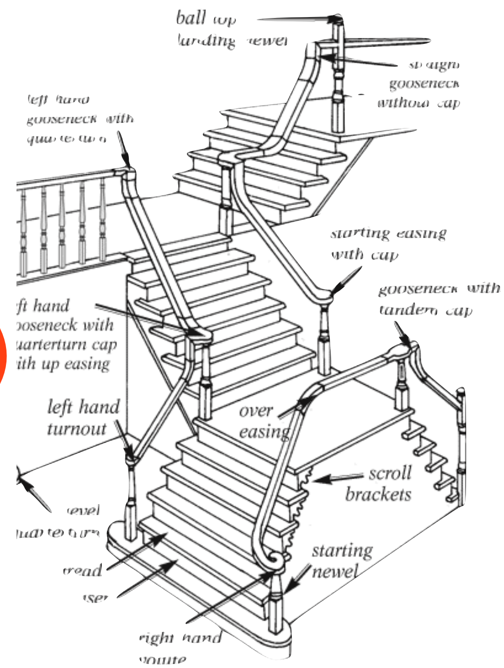


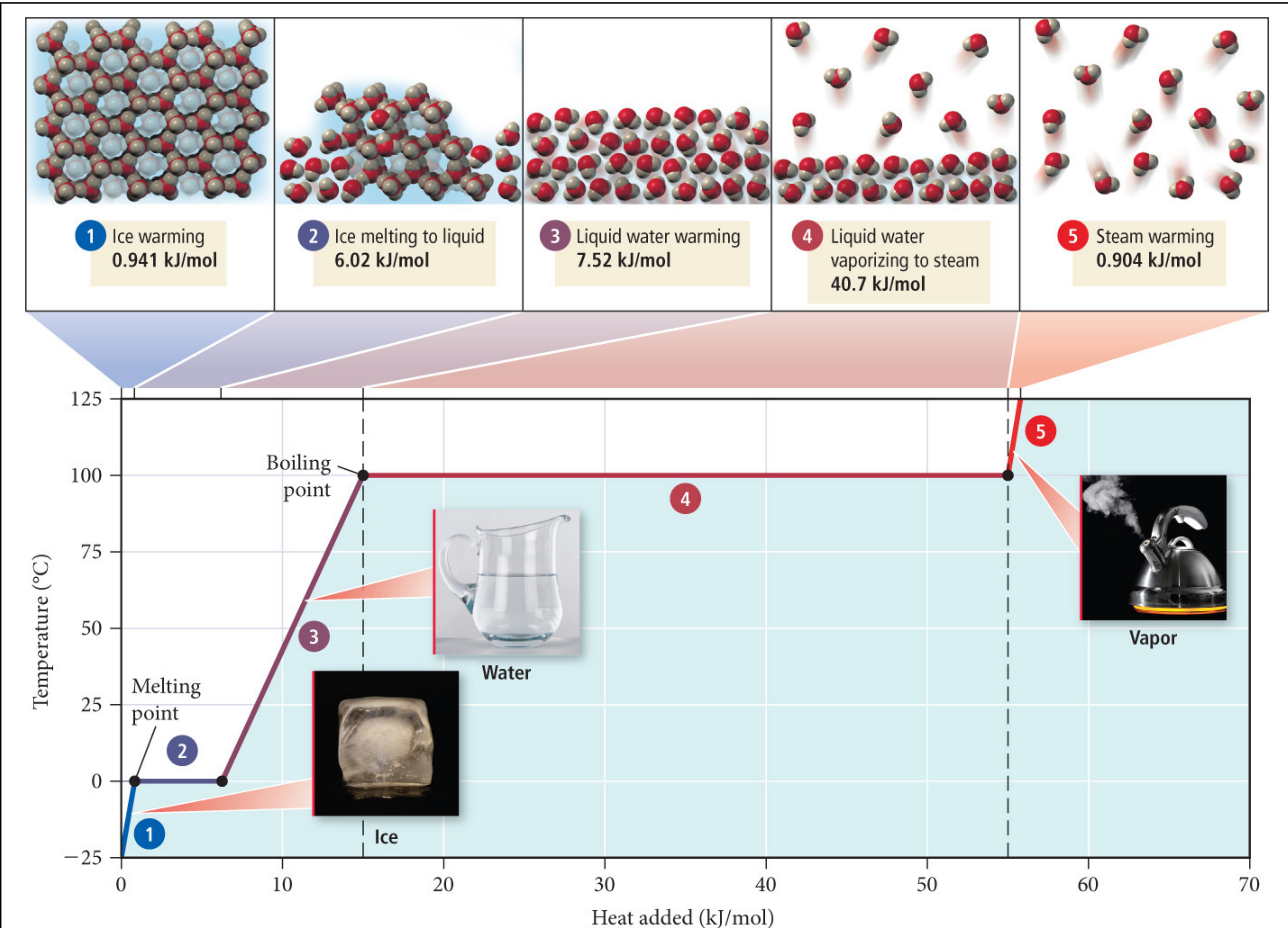
Heat Effects Matter

- Heat (q) has two principle effects on matter.
  - It can cause a change in temperature.
  - It can cause matter to change state.
- Which effect occurs, *what the heat does*, depends on where you are relative to threshold temperatures.
  - Melting Point
  - Boiling Point

Temperature	Effect on Water
Above 100˚C	Gas Changes Temperature
At exactly 100˚C	Liquid Converts to Gas
Between 0˚- 100˚C	Liquid Changes Temperature
At exactly 0˚C	Solid Converts to Liquid
Below 0˚C	Solid Changes Temperature

Heating Water and Boiling Water are very different processes!





1 Ice warming
0.941 kJ/mol
2 Ice melting to liquid
6.02 kJ/mol
3 Liquid water warming
7.52 kJ/mol
4 Liquid water
vaporizing to steam
40.7 kJ/mol
5 Steam warming
0.904 kJ/mol
Temperature (°C)
125
100
75
50
25
0
−25
Boiling point
Melting point
Water
Ice
Vapor
0
10
20
30
40
50
60
70
Heat added (kJ/mol)

Heat Effects Matter

- Heat (q) has two principle effects on matter.
  - It can cause a change in temperature.
  - It can cause matter to change state.
- Which effect occurs, *what the heat does*, depends on where you are relative to threshold temperatures.
    - Melting Point
    - Boiling Point
- Heat Capacity allows you calculate the heat required to change temperature.
- Heat of Fusion allows you to calculate the heat required to change state.
- Both equations require you to consider the quantity of the substance.
  - The quantity can represented as
    - moles (n) or mass (m).
- In either case you need the appropriate conversion factor.
  - C_s (specific heat for mass)
  - C_m (molar heat capacity for moles)
  - ΔH_s (specific heat of fusion for mass)
  - ΔH_m (molar heat of fusion for moles)

Temperature	Effect on Water
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At exactly 0˚C	Solid Converts to Liquid
Below 0˚C	Solid Changes Temperature

Heat Capacity

$$q = C\,\Delta T$$

$$C = C_s \cdot m$$
$$C = C_m \cdot n$$

Heat of Fusion

$$q = \Delta H$$

$$\Delta H = \Delta H_s \cdot m$$
$$\Delta H = \Delta H_m \cdot n$$

Heat Effects Matter

- Heat (q) has two principle effects on matter.
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  - It can cause matter to change state.
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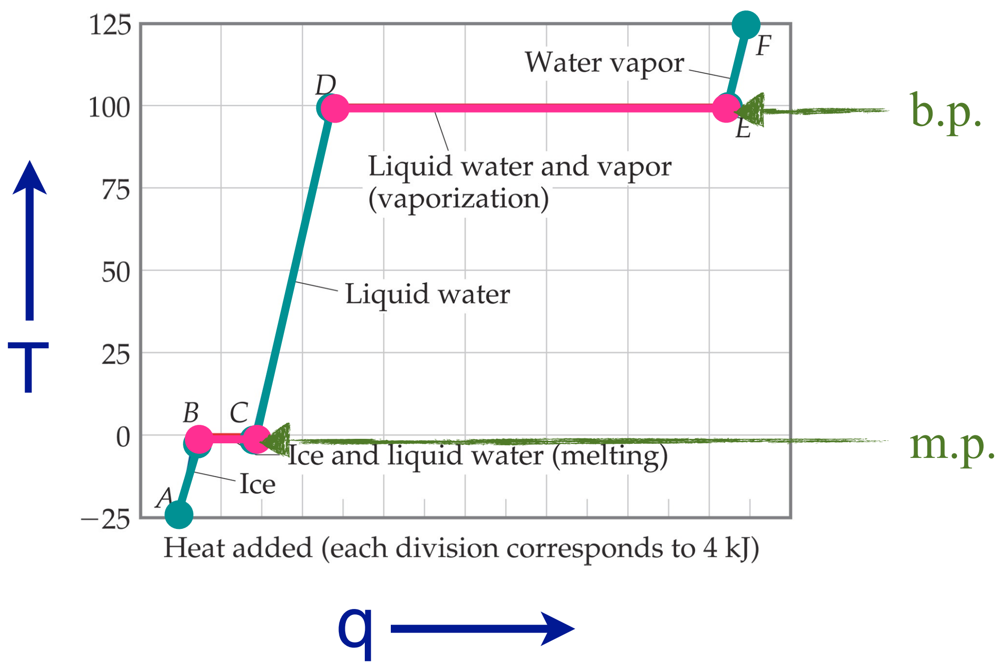
At exactly 0˚C	Solid Converts to Liquid
Below 0˚C	Solid Changes Temperature



Heat Capacity

$$q = C\,\Delta T$$

$$C = C_s \cdot m$$
$$C = C_m \cdot n$$

Heat of Fusion

$$q = \Delta H$$

$$\Delta H = \Delta H_s \cdot m$$
$$\Delta H = \Delta H_m \cdot n$$

Problem:

What is the heat in kilojoules required to heat 25.0 grams of ice from -25.0 ˚C to -5.0 ˚C?

Temperature	Effect on Water
Above 100˚C	Gas Changes Temperature
At exactly 100˚C	Liquid Converts to Gas
Between 0˚- 100˚C	Liquid Changes Temperature
At exactly 0˚C	Solid Converts to Liquid
Below 0˚C	Solid Changes Temperature

heat of fusion of water = 334 J/g
heat of vaporization of water = 2257 J/g
specific heat of ice = 2.09 J/g·°C
specific heat of water = 4.18 J/g·°C
specific heat of water vapor = 1.99 J/g·°C

Specific Heat

$$q = C_s \cdot m \cdot \Delta T \;;\; J \rightarrow kJ$$

m = 25.0 grams

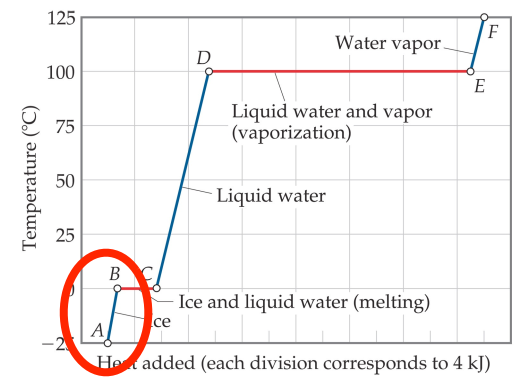
q = heat required

$C_{s\ ice}$ = 2.09 J/g˚C

T_F = -25.0 ˚C
T_I = -5.0 ˚C

ΔT = 20.0 ˚C

(ΔT is positive because you should be doing this in Kelvin)



$$q = 2.09\ \frac{J}{g^{\circ}C} \times 25.0\ g \times 20^{\circ}C = 1045\ J$$

$$q = 1045\ J \times \frac{1kJ}{1000J} = \boxed{1.05\ kJ}$$

Problem:

What is the heat in kilojoules required to melt 25.0 grams of ice?

Temperature	Effect on Water
Above 100˚C	Gas Changes Temperature
At exactly 100˚C	Liquid Converts to Gas
Between 0˚ - 100˚C	Liquid Changes Temperature
At exactly 0˚C	Solid Converts to Liquid
Below 0˚C	Solid Changes Temperature

heat of fusion of water = 334 J/g
heat of vaporization of water = 2257 J/g
specific heat of ice = 2.09 J/g·°C
specific heat of water = 4.18 J/g·°C
specific heat of water vapor = 1.99 J/g·°C

Heat of Fusion

$$q = \Delta H_s \bullet m \text{ ; J} \rightarrow \text{kJ}$$

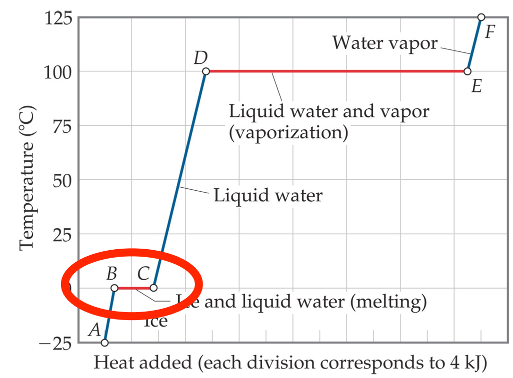
m = 25.0 grams

q = heat required

ΔH_{fus} = 334 J/g

$$q = 334 \frac{J}{g} \times 25.0 \text{ g}$$

$$q = 8350 \text{ J} \times \frac{1 \text{kJ}}{1000 \text{J}} = \boxed{8.35 \text{kJ}}$$



Problem:

What is the heat in kilojoules required to heat 25.0 grams of water from 0.0˚C to 73.8 ˚C?

Temperature	Effect on Water
Above 100˚C	Gas Changes Temperature
At exactly 100˚C	Liquid Converts to Gas
Between 0˚- 100˚C	Liquid Changes Temperature
At exactly 0˚C	Solid Converts to Liquid
Below 0˚C	Solid Changes Temperature

heat of fusion of water = 334 J/g
heat of vaporization of water = 2257 J/g
specific heat of ice = 2.09 J/g·°C
specific heat of water = 4.18 J/g·°C
specific heat of water vapor = 1.99 J/g·°C

Specific Heat

$$q = C_s \cdot m \cdot \Delta T \text{ ; } J \rightarrow kJ$$

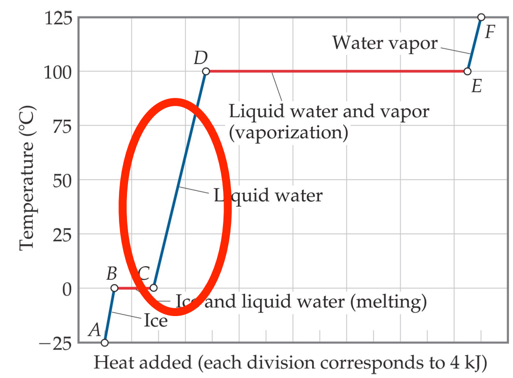
$m = 25.0$ grams

q = heat required

$C_{s\ water} = 4.18$ J/g˚C

$T_F = 73.8$ ˚C
$T_I = 0.0$ ˚C

$\Delta T = 73.8$ ˚C



$$q = 4.18\ \frac{J}{g^\circ C} \times 25.0\ g \times 73.8^\circ C = 7712.1\ J$$

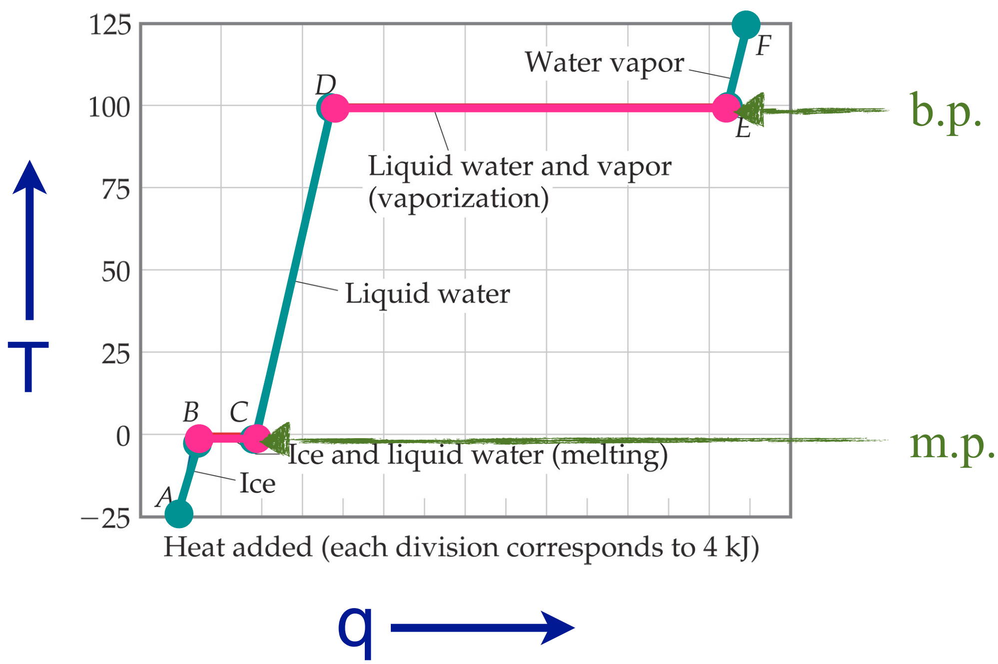
$$q = 7712.1\ J \times \frac{1kJ}{1000J} = \boxed{7.71\ kJ}$$

Heat Effects Matter

- Heat (q) has two principle effects on matter.
  - It can cause a change in temperature.
  - It can cause matter to change state.
- Which effect occurs, *what the heat does*, depends on where you are relative to threshold temperatures.
- Heat is a state function. It's additive.

Temperature	Effect on Water
Above 100˚C	Gas Changes Temperature
At exactly 100˚C	Liquid Converts to Gas
Between 0˚- 100˚C	Liquid Changes Temperature
At exactly 0˚C	Solid Converts to Liquid
Below 0˚C	Solid Changes Temperature

$$q_{B\text{-}D} = q_{B\text{-}C} + q_{C\text{-}D}$$



Heat Capacity

$$q = C\,\Delta T$$

$$C = C_s \cdot m$$

$$C = C_m \cdot n$$

Heat of Fusion

$$q = \Delta H$$

$$\Delta H = \Delta H_s \cdot m$$

$$\Delta H = \Delta H_m \cdot n$$

Problem:

What is the heat in kilojoules required to heat 25.0 grams of ice from 0.0˚C to 73.8 ˚C?

Temperature	Effect on Water
Above 100˚C	Gas Changes Temperature
At exactly 100˚C	Liquid Converts to Gas
Between 0˚ - 100˚C	Liquid Changes Temperature
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specific heat of ice = 2.09 J/g·°C
specific heat of water = 4.18 J/g·°C
specific heat of water vapor = 1.99 J/g·°C

Heat of Fusion $q_1 = \Delta H_s \cdot m$

Specific Heat $q_2 = C_s \cdot m \cdot \Delta T$

$q = q_1 + q_2$; J → kJ

m = 25.0 grams

q = heat required

ΔH_{fus} = 334 J/g

$C_{s\ water}$ = 4.18 J/g˚C

T_F = 73.8 ˚C

T_I = 0.0 ˚C

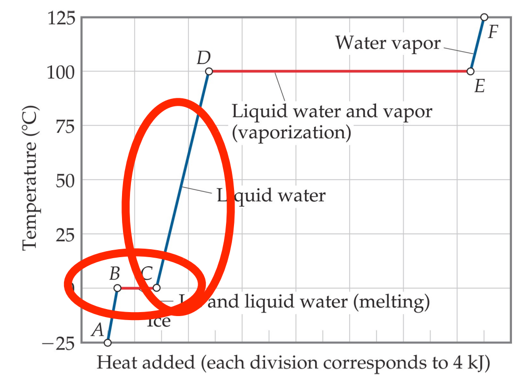
ΔT = 73.8 ˚C

$$q_1 = 334\ \frac{J}{g} \times 25.0\ g = 8350\ J$$

$$q_2 = 4.18\ \frac{J}{g^\circ C} \times 25.0\ g \times 73.8^\circ C = 7712.1\ J$$

$$q = q_1 + q_2 = 8350\ J + 7712.1\ J = 1{,}6062.1\ J$$

$$q = 1{,}6062.1\ J \times \frac{1kJ}{1000J} = 16.06\ kJ$$



8350
7710
16060

A Steam Engine

Water at 45.5° is heated to convert it entirely to steam at exactly 100°C. How much heat is needed to convert 23.2 grams?

Temperature	Effect on Water
Above 100°C	Gas Changes Temperature
At exactly 100°C	Liquid Converts to Gas
Between 0° - 100°C	Liquid Changes Temperature
At exactly 0°C	Solid Converts to Liquid
Below 0°C	Solid Changes Temperature

heat of fusion of water = 334 J/g
heat of vaporization of water = 2257 J/g
specific heat of ice = 2.09 J/g·°C
specific heat of water = 4.18 J/g·°C
specific heat of water vapor = 1.99 J/g·°C

45.5–100° – Liquid water Heats

$T_f = 100.0$
$T_I = 45.5\,°C$
$\Delta T\ 54.5\,°C$
$m = 23.2\,g$
$C_s = 4.18\ J/g°C$

$q_1 = C_s \cdot m \cdot \Delta T$
$= 4.18\ J/g°C \cdot 23.2\,g \cdot 54.5\,°C$
$= 5{,}285.192\ J$
$= 5{,}290\ J$

Total Heat = $q_1 + q_2$

$= 57{,}700\ J$

100° – Liquid water changes to Steam

$m = 23.2\,g$
$\Delta H_{vap} = 2257\ J/g$

$q_2 = \Delta H_{vap} \cdot mass$
$= 2257\ J/g \cdot 23.2\,g$
$= 52{,}362.40\ J$
$= 52{,}400\ J$

5290
+ 52400
57690

Melting Ice

A 35.0 g block of metal at 80.0˚C is added to a mixture of 100.0 g water and 15.0 g of ice in an isolated container. All the ice melted and the temperature in the container rose to 10.0 ˚C.

How much heat came out of the metal?

Temperature	Effect on Water
Above 100˚C	Gas Changes Temperature
At exactly 100˚C	Liquid Converts to Gas
Between 0˚- 100˚C	Liquid Changes Temperature
At exactly 0˚C	Solid Converts to Liquid
Below 0˚C	Solid Changes Temperature

heat of fusion of water = 334 J/g
heat of vaporization of water = 2257 J/g
specific heat of ice = 2.09 J/g·°C
specific heat of water = 4.18 J/g·°C
specific heat of water vapor = 1.99 J/g·°C

Heat out of metal = Heat into ice/water
Ice Water is 0°C (exactly)

Melts 15.0 g ice at 0°C

mass = 15.0 g
ΔH_{fus} = 334 J/g

$q_1 = \Delta H_{fus} \cdot mass$
$= 334\ J/g \cdot 15.0\ g$
$= 5{,}010.00$
$= 5{,}010\ J$

5,010
4810
9820

$q_F = q_1 + q_2$
$= 9{,}820\ J$

$= 9.82\ kJ$

HEATING WATER 0°C → 10.0°C

$T_F = 10.0°C$
$T_I = 0°C$ (exact)
$\Delta T = 10.0°C$

$C_s = 4.18\ J/g°C$

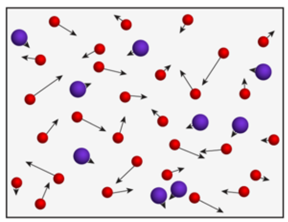
mass = 100.0 + 15.0 g
= 115.0 g

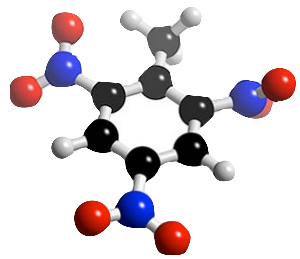
$q_2 = C_s \cdot m \cdot \Delta T$
$= 4.18\ J/g°C \cdot 115.0\ g \cdot 10.0°C$
$= 4{,}810\ J$

Ch03

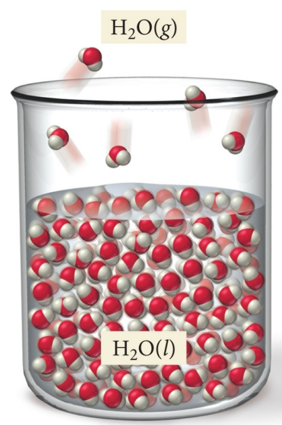
Energy Affecting Matter

- Energy
  - Defined
  - Kinetic vs Potential
    - Thermal vs Chemical
  - Units of Energy
- Energy Content
  - Energy value
  - Total energy
- Temperature
  - Temperature scales
    - Celsius, Fahrenheit, Kelvin
    - Converting
  - Measurement Types
    - Size, Position, Change
    - Finding ΔX





- Energy Affecting Matter
  - Heat
    - Heat Capacity
    - Example Calculations
  - State
    - Critical Temperatures
    - Example Calculations



Questions?