

## Announcements – 10 Nov 2009

1. Exam 4 not that far away...
  - a. Starts a week from Thursday! Yikes!
  - b. We'll have another evening TA review...I'll send out a doodle.com survey very soon

Colton - Lecture 21 - pg 1

## Demos we didn't do last time:

- Two thermometers
- Bimetallic strip
- Ball & washer
- Liquid nitrogen
  - Rubber nail
  - Two balloons
  - Pipe "balloon pop"

**We worked this problem:** Use the ideal gas law to determine the density of air at 1 atm and 300 K (80° F). ( $MM_{\text{air}} = 29 \text{ g/mol}$ )

Start with:  $PV = nRT$

Plug in:  $n = \frac{m}{MM}$

You get:  $PV = \left(\frac{m}{MM}\right)RT$

Solve for m/V:  $\frac{m}{V} = \frac{P(MM)}{RT}$

With the given numbers:

$$\frac{m}{V} = \frac{(1.01 \cdot 10^5 \text{ Pa})(0.029 \text{ kg/mol})}{(8.31 \text{ J/mol} \cdot \text{K})(300 \text{ K})} = 1.175 \text{ kg/m}^3$$

Colton - Lecture 21 - pg 2

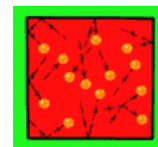
**Difficult Worked Problem:** A hot air balloon wants to lift off on an 80° F day. The balloon fabric and basket weight 200 kg, and there are four 80 kg passengers. The balloon is spherical, with an 8 m radius. How hot do they have to get the air inside the balloon? *Hint:* Do not neglect the weight of the hot air inside the balloon!

**Plan:** (a) figure out the maximum mass of hot air, (b) then the density of the hot air, then (c) figure out what temperature gives that density

Answers: 2000.0 kg; 0.9325 kg/m<sup>3</sup>; 378 K

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## Molecular view



### Equipartition Theorem

Only briefly mentioned in your book! And not by name!  
(see page 390, Section 12.2 in 8<sup>th</sup> edition)

“The total kinetic energy of a system is shared equally among all of its independent parts, on the average, once the system has reached thermal equilibrium.”

Specifically, each “degree of freedom”,  
of each molecule, has “thermal energy” of: \_\_\_\_\_

“independent parts”: larger for molecules that can

- rotate
- vibrate

(requires more than one atom)

→ **such molecules have more “internal energy”**

*Translational* kinetic energy: Three independent directions

**Result:**  $KE_{\text{ave}} = \frac{1}{2}mv_{\text{ave}}^2 = \frac{3}{2}k_B T$

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**Clicker quiz:** In air, the molecular mass of oxygen molecules is 32 g/mol; the molecular mass of nitrogen molecules is 28 g/mol. Which molecules are traveling faster on average?

- c. Oxygen
- d. Nitrogen
- e. Same speed

**Demo:** molecular speed

**Worked Problem:** How fast are the oxygen molecules traveling if the air is 300 K?

Answer: 483.46 m/s (= 1081 mph!)

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## Molecular View of Pressure

**Pressure:** Comes from collision forces of molecules hitting wall

**Related problem:** You throw baseballs (mass 145 g) at a wall (area 9 m<sup>2</sup>), at a speed of 85 mph (38 m/s). The collisions are elastic, and last for 0.05 seconds. (This is the time the ball is in contact with the wall.) A baseball hits the wall every 0.5 seconds.

(a) How much force is generated by each hit? (Use impulse)

(b) How much force is there, on average?

(c) How much overall pressure is generated by the balls?

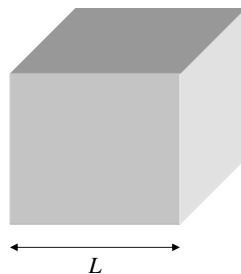
Answers: 220.4 N; 22.04 N; 2.449 Pa

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**The actual problem:** A cube filled with gas (focus on x-direction for now)

Molecules (mass  $m$ ) hit the right wall, at a speed of  $v_x$ . Elastic collisions.

(a) How much force is generated by each hit? (Use impulse)



(b) How much force is there, on average?

(c) How much pressure is generated by the molecules?

(d) Expand to  $N$  molecules, and 3 dimensions ( $v_x = v_y = v_z$ ).  $P = ?$

Answers:  $\frac{2mv_x}{\text{time of collision}}$ ;  $\frac{mv_x^2}{L}$ ;  $\frac{mv_x^2}{V}$ ;  $\frac{Nm(\frac{1}{3}v^2)}{V}$

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How does this relate to the Equipartition Theorem result?

$$\text{Recall: } KE_{ave} = \frac{1}{2}mv_{ave}^2 = \frac{3}{2}k_B T$$

$$v_{ave} = \sqrt{\frac{3k_B T}{m}}$$

Use Ideal Gas Law...

**Take home message:**

**From warmup:** An ideal gas has a mixture of heavy and light molecules at the same temperature. The molecules with the most KE are...

- a. heavy
- b. light
- c. same

**Demos:** kinetic theory machine

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## Heat

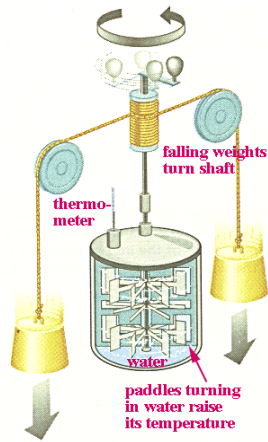
Heat is random kinetic energy!

Symbol:  $Q$

Units: Joules

"Mechanical equivalent of heat":

James Joule 1849



calories and Calories:

1 calorie = 4.186 J

Food calorie: 1 Cal

= a *kilocalorie*

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## "Specific heat"

How much does  $T$  rise when heat energy is added?

- temperature rise is proportional to heat added
- the more mass... the less the temperature rises
- material dependent

$$Q = mc\Delta T$$

$c$  = "specific heat"  
(closely related to  
"heat capacity")

$mc$  sometimes called  
"thermal mass"

TABLE 11.1

Specific Heats of Some  
Materials at Atmospheric  
Pressure

| Substance | J/kg · °C | cal/g · °C |
|-----------|-----------|------------|
| Aluminum  | 900       | 0.215      |
| Beryllium | 1 820     | 0.436      |
| Cadmium   | 230       | 0.055      |
| Copper    | 387       | 0.0924     |
| Germanium | 322       | 0.077      |
| Glass     | 837       | 0.200      |
| Gold      | 129       | 0.0308     |
| Ice       | 2 090     | 0.500      |
| Iron      | 448       | 0.107      |
| Lead      | 128       | 0.0305     |
| Mercury   | 138       | 0.033      |
| Silicon   | 703       | 0.168      |
| Silver    | 234       | 0.056      |
| Steam     | 2 010     | 0.480      |
| Water     | 4 186     | 1.00       |

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**Clicker quiz:** If you add 5 J of heat to a mass of water, and 5 J of heat to the same mass of copper, which one increases the most in temperature?

- Water
- Copper
- Same

**From warmup:** The fact that desert sand is very hot in the day and very cold at night is evidence that sand has a:

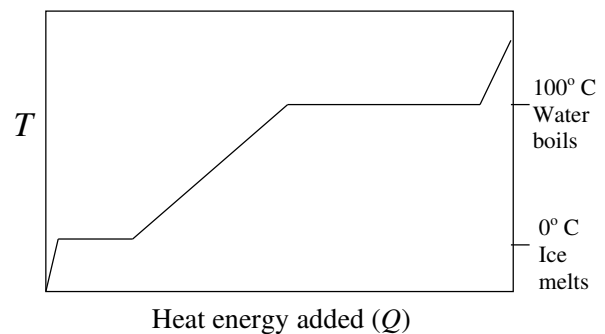
- low specific heat
- high specific heat

## Next topic: Phase Changes

**From warmup:** Thermal energy that is used to melt or freeze something is called:

- latent heat
- specific heat
- thermal mass

## Phase Changes



During phase change, no  $T$  increase

→ but heat still needed to complete the phase change

→ both phases co-exist

$$Q = mL$$

$L$  depends on

- Material
- Type of phase change (i.e. solid-liquid, liquid-gas, or other)

Water:

$$L_{\text{melting/freezing}} = 3.33 \times 10^5 \text{ J/kg}$$

$$L_{\text{boiling/condensing}} = 2.26 \times 10^6 \text{ J/kg}$$

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Colton - Lecture 21 - pg 12

**Clicker quiz:** If you want to melt a cube of ice that's initially at  $-40^{\circ}\text{C}$ , which part takes the most energy?

- Raising the temperature
- Converting from solid to liquid phase
- Same

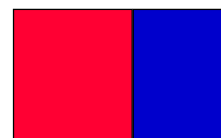


**From warmup:** Ralph's professor stated "If you add an ice cube to a glass of water, the temperature of the water does not necessarily decrease." That seems bizarre to him, because ice is obviously used to cool down water! Can you help him understand what his professor may have been talking about?

**Answer from the class:**

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## Calorimetry



Conservation of energy:

$$Q_{\text{gained by cold objects}} = Q_{\text{lost by hot objects}}$$

(assuming no heat flow to outside)

→ On both sides of equation use only *positive* quantities

→ May need to include melting and boiling:  $mL$  terms

**Worked Problem:** 500 g iron at  $300^{\circ}\text{C}$  added to 100 g of water at  $30^{\circ}\text{C}$ . How much water boils away?

Start with:  $Q_{\text{gained by water}} = Q_{\text{lost by iron}}$

Answer: 6.86 g

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**Worked Problem (if time):** (a) 5 g of hot iron at  $300^{\circ}\text{C}$  is added to 100 g of water at  $30^{\circ}\text{C}$ . What is the final temperature? (b) Repeat, but with 500 g iron

Set up for both:  $Q_{\text{gained by water}} = Q_{\text{lost by iron}}$

(a)

(b)

Answers:  $31.44^{\circ}\text{C}$ ; 124.1 (not real answer),  $-395.3^{\circ}\text{C}$  (not real answer),  $100^{\circ}\text{C}$

Colton - Lecture 21 - pg 15

Colton - Lecture 21 - pg 16