

## Phy 212: General Physics II

### Chapter 18: Temperature, Heat & the 1<sup>st</sup> Law of Thermodynamics Lecture Notes

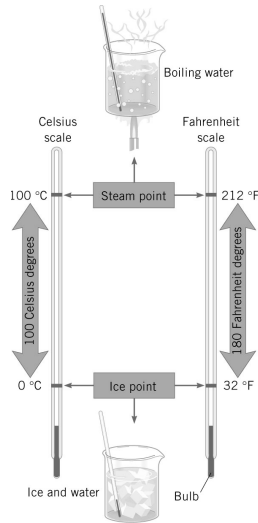
### What is Temperature?

1. Temperature (T) is a measure of how “hot” or “cold” something is
2. Temperature is not energy nor is it heat (more on this later)
3. Temperature is a measure of the motion (vibration or translation) of the atoms/molecules that make-up an object
  - a. The greater the motion/vibration the greater the T
  - b. The smaller the motion/vibration the lower the T
4. Temperature is measured in 3 common scales:
  - a. Fahrenheit (USCS), °F
  - b. Celsius or Centigrade (Metric), °C
  - c. Kelvin (SI), K{note: units of Kelvin are not degrees K (or °K), just K}
5. In Kelvin units, temperature of a system is proportional to the average energy of the particles in the system:

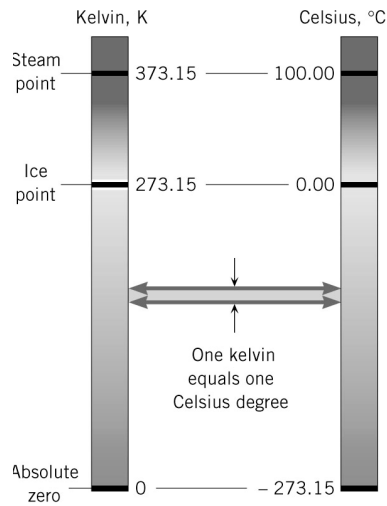
$$E_{\text{avg}} \propto T_K$$

# The Common Temperature Scales

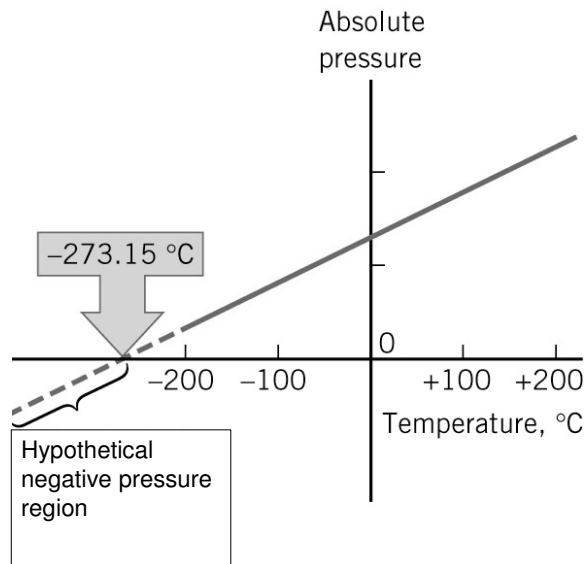
## Fahrenheit & Celsius



## Kelvin & Celsius



## Absolute Zero (0K)



## Linear Thermal Expansion

- Most solid materials expand or contract depending on their temperature (called linear thermal expansion):
  - As an object's T increases it expands
  - As an object's T decreases it contracts
- The relation between length and temperature:

$$\Delta L = \alpha L_o \Delta T \quad \text{or} \quad \frac{\Delta L}{\Delta T} = \alpha L_o$$

**where:**

$\alpha$  = the coeff. of linear thermal expansion

$L_o$  = the original length of the object

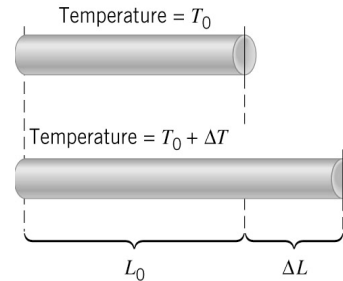
$\Delta T$  = temperature change in °C or K

**Note:**  $L_f = L_o(1 + \alpha \Delta T)$  or

$$\frac{dL}{dT} = \lim_{\Delta T \rightarrow 0} \left( \frac{\Delta L}{\Delta T} \right) = \alpha L_o$$

$$\frac{dL}{L} = \alpha dT \Rightarrow \int_{L_o}^{L_f} \frac{dL}{L} = \alpha \int_{T_o}^{T_f} dT \Rightarrow \ln \left( \frac{L_f}{L_o} \right) = \alpha \Delta T \Rightarrow L_f = L_o e^{\alpha \Delta T}$$

**Old saying:** "you can tell the temperature by the length of the cat..."



## Volume Thermal Expansion

- Objects will expand in volume as well as in length due to changes in temperature
- Volume thermal expansion is analogous to linear thermal expansion:

$$\Delta V = \beta V_o \Delta T \quad \text{or} \quad \frac{\Delta V}{\Delta T} = \beta V_o \Rightarrow V_f = V_o e^{\beta \Delta T}$$

**where:**

$\beta$  = the coefficient of volume thermal expansion

$V_o$  = the original volume of the object

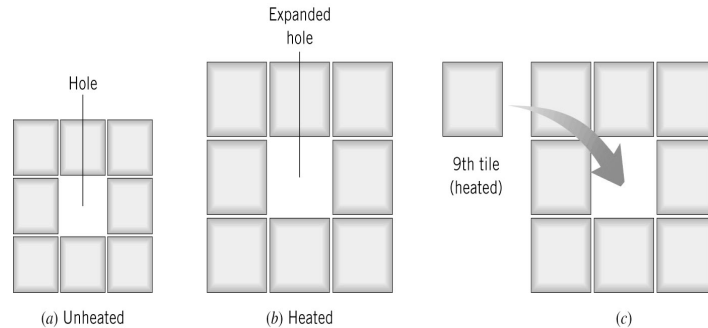
$\Delta T$  = temperature change in °C or K units

- The coefficient,  $\beta$ , is called the coefficient of volume thermal expansion & is related to the coefficient of linear thermal expansion ( $\alpha$ )

$$\beta = 3\alpha$$

- Not only do solid objects expand/contract with temperature but so do holes within them!

## Consider a “square” hole in a solid body



1. Think of the solid body as a collection of connected solid cubes
2. The small cubes each expand with increased temperature, which expands the hole in the middle
3. If a block (same size as the original hole) is heated to the same (higher) temperature it will fit perfectly in the hole

***The hole expands/contracts as if it were a solid body of that material!***

## Heat & Internal Energy

1. Heat is energy that flows from a high temperature object to a lower temperature object
  - a. When something absorbs heat its internal energy (or the energy of its atoms/molecules) increases
  - b. When something releases heat its internal energy decreases
2. The SI units of heat energy are joules (J)
3. Two things occur when an object absorbs or releases heat energy (Q):
  - a. The temperature will change (which is why they expand/contract, due to changes in molecular motion)
  - b. The object (or part of it) will change phase (solid, liquid, gas)

## Heat & Temperature Change

1. When an solid (or liquid) of mass,  $m$ , absorbs and no phase change occurs:

$$Q = c \cdot m \cdot \Delta T \text{ \{Black's Heat Equation\}}$$

$Q$  is heat absorbed

$\Delta T$  is the temperature change

$c$  is the specific heat capacity

2. Specific Heat Capacity:

- a. The physical property that relates how energy absorbed reflects changes in temperature for a given mass of a substance  
*In SI terms: the amount of energy required to change the temperature of 1 kg of a substance by 1°C (or 1 K)*
- b. SI units for specific heat capacity are J/(kg·K)
- c. Metals tend to have low specific heat capacities (which is one reason they make great cooking vessels)
- d. Non-metal substances tend to have higher specific heat capacities
- e. Water has an unusually high specific heat capacity

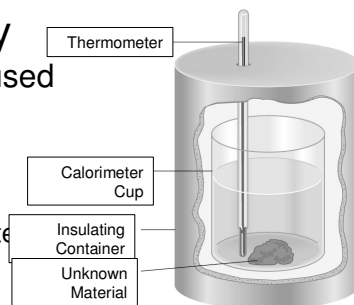
## The Mechanical Equivalent of Heat

1. Before it was established that heat was indeed energy, heat was traditionally measured in units of calories (cal):  
1 calorie = amount of heat required to raise the T of 1 g of water by 1 °C
2. James Prescott Joule established that heat energy is the equivalent of mechanical energy: 1 cal = 4.186 J
3. In nutrition, we often invoke the kilocalorie (or nutritional calorie):  
 $1 \text{ kcal} = 1 \text{ Cal} = 1000 \text{ cal} = 4186 \text{ J}$
4. The "typical" person needs to eat ~2000 kcal daily (or  $2 \times 10^6$  cal of food energy)
  - a. How much energy in J and kcal is required to raise the body temperature of a "typical" person ( $m=50 \text{ kg}$ ) from 20°C to 31°C? Assume the body is essentially all water.
  - b. A typical candy has 250 kcal. What is the limit of vertical displacement you could climb using only this energy?

## Calorimetry

Calorimetry is an experimental method used to measure unknown specific heat capacity that is based on:

- The principle of conservation of energy
- The known specific heat capacity for water (1000 cal/kg·°C or 4186 J/kg·°C)



The Process:

- A heated object is placed into a thermally isolated container containing a known amount of water

$$Q_{\text{net}} = Q_{\text{gained by water}} + Q_{\text{gained by object}} = 0$$

- Object and liquid reach thermal equilibrium:

$$Q_{\text{gained by water}} = -Q_{\text{gained by object}} \{ = Q_{\text{lost by object}} \}$$

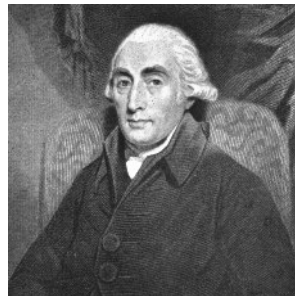
- The final temperature of the object/water is used to determine the specific heat of the object

## Typical Calorimetry Problems

- An unknown object of 0.10 kg mass at initial temperature of 90.0 °C is placed in a “calorimeter” containing a 0.10 kg of water at 15.0 °C.
  - If the final temperature of the water/object is 22.5 °C, what is the specific heat capacity of the object?
  - What is the substance made of?
- An 0.15 kg sample of gold ( $c_{\text{gold}} = 129 \text{ J/kg}\cdot\text{K}$ ) is heated to 95.0 °C then placed in a thermally isolated container containing 0.10 kg of water at 15.0 °C.
  - What is the ratio of  $\Delta T_{\text{gold}}/\Delta T_{\text{water}}$ ?
  - What is the final temperature of the object/water?

## Joseph Black (1728-1799)

1. English chemist who developed the calorimetry technique
2. He (re-)discovered carbon dioxide ("fixed air")
3. He proposed a theory of heat that did not reference phlogiston, including:
  - a. The heat equation
  - b. Specific heat capacity & Latent heat
4. He widely renowned as a great teacher:



*"Many were induced, by the report of his students, to attend his courses, without having any particular relish for chemical knowledge."*

## Heat & Phase Change

1. When a substance absorbs or loses heat energy its temperature will change until:
  - a. the substance reaches its "critical" temperature it will no longer change temperature
  - b. The gain/loss of additional heat energy will also result in the phase transformation of the matter from one phase to another:

Solid → liquid → gas

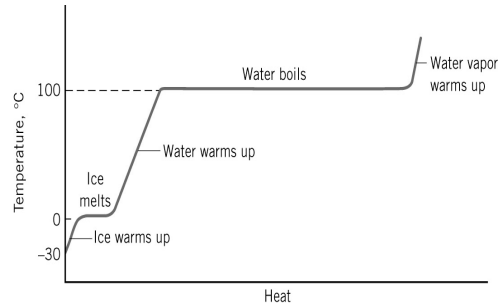
2. The relationship that describes heat energy (Q) gained/lost to mass of substance (m) that undergoes a phase change is:

$$Q = m \cdot L$$

*L is the latent heat for the phase transition* (the SI units are J/kg)

3. The latent heat is a physical property that describes how much energy is required to transform the mass of a substance from one phase to another (e.g. latent heat of fusion or vaporization, etc.)

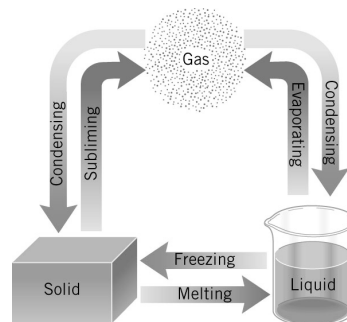
## Heat vs. Temperature Graph (for water)



### Question:

Why do you feel so “cold” when you step out of the shower soaking wet?

### Examples of Phase Change



## Work: A Deeper Perspective

1. Work can be performed on a system
  - a. Work performed on a system increases the internal energy of the system & the net energy change for the system will be positive
2. Work can be performed by a system
  - a. Work performed on by system decreases the internal energy of the system & the net energy change for the system will be negative
3. The work performed on a closed (fluid) system is given by:

$$W = \int dW = \int_{V_i}^{V_f} p dV$$

where:  $p$  = fluid pressure and  $V$  = fluid volume



## First Law of Thermodynamics

The 1<sup>st</sup> Law of Thermodynamics is a statement of the Law of Energy Conservation for a system

1. The state of the system is defined by its internal energy,  $E_{\text{int}}$
2. Factors that change a system's internal energy
  - a. When a system absorbs heat ( $Q$ ),  $E_{\text{int}}$  increases
  - b. When a system performs work ( $W$ ),  $E_{\text{int}}$  decreases
  - c. When work ( $-W$ ) is performed on the system,  $E_{\text{int}}$  increases
3. Expressed as an equation,

$$\Delta E_{\text{int}} = E_{\text{int},f} - E_{\text{int},i} = Q - W \quad \text{or} \quad dE_{\text{int}} = dQ - dW$$

Where:

$Q$  is the heat energy absorbed by the system

$W$  is the work performed by the system

$E_{\text{int},f}$  and  $E_{\text{int},i}$  are the system's final & initial internal energy states

**Note:** It is assumed that the kinetic & potential energies, respectively, of the system as a whole remain constant:  $\Delta K_{\text{system}} = 0$  &  $\Delta U_{\text{system}} = 0$

## Applications of the 1<sup>st</sup> Law of Thermodynamics

The 1<sup>st</sup> Law of Thermodynamics can be applied to several special cases:

1. *Adiabatic: a process in which there is no heat exchange with the outside environment*  $Q = 0 \Rightarrow \Delta E_{\text{int}} = -W$
2. *Constant-volume: When the volume of a system is fixed, the system cannot perform work*  $W = 0 \Rightarrow \Delta E_{\text{int}} = Q$
3. *Closed cycle: A sequence of processes after which the internal energy of the system returns to its initial state*  
 $Q = W \Rightarrow \Delta E_{\text{int}} = 0$
4. *Free expansions: Adiabatic processes in which no work is performed on or by the system*  $Q = W \Rightarrow \Delta E_{\text{int}} = 0$

## Summary of Thermal Transfer Processes

### 1. Conduction

- Thermal energy (heat) that is transferred directly through a material or across an interface (in contact) between 2 materials

### 2. Convection

- Thermal energy that is carried from place to place by the bulk movement of a fluid, i.e. liquid or gas

### 3. Radiation

- Thermal energy that is radiantly transferred by means of electromagnetic waves, i.e. radiant energy, light, etc.

## Conduction

- Heat flow due to collisions between neighboring atoms (a sort of “domino effect”)
- The rate of conductive heat flow ( $Q/\Delta t$ ) is:
  - proportional to temperature difference between 2 regions in a conducting pathway material ( $\Delta T$ )
  - proportional to cross sectional area of material ( $A$ )
  - proportional to ability of material to conduct heat ( $k$ ) {called thermal conductivity}
  - inversely proportional to length of conducting pathway ( $L$ )
- Combining all of these elements forms *Fourier's Heat Equation* (1609):

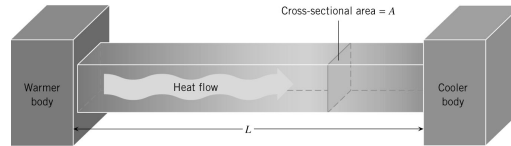
$$\text{Power} = \frac{Q}{\Delta t} = \left( \frac{kA}{L} \right) \Delta T \quad \{\text{Rate of Heat Flow in W}\}$$

or

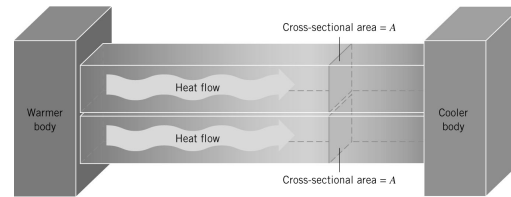
$$\text{Power} = \lim_{\Delta t \rightarrow 0} \left( \frac{Q}{\Delta t} \right) = \frac{dQ}{dt} = \frac{kA}{L} dT$$

## Conductive Heat Flow

1. Conduction depends on temperature difference between 2 regions & how far apart those regions are separated



2. Increasing the cross-sectional area increases amount of heat that will flow in a given time



3. Of course, the relative ability to conduct heat is an intrinsic property of different materials !!

## Jean Baptiste Fourier (1768-1830)

1. French mathematician
2. Originally wanted to become a priest but drawn to mathematics
3. Contemporary of Laplace, Lagrange, Biot & Poisson
4. Served as Precept under Napoleon
5. Worked on:
  1. mathematics
  2. heat conduction

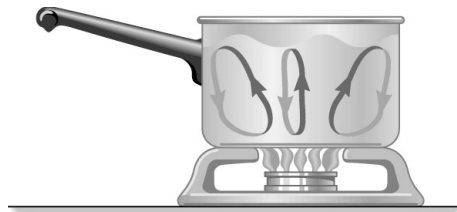


*Yesterday was my 21st birthday, at that age Newton and Pascal had already acquired many claims to immortality.*

## Convection

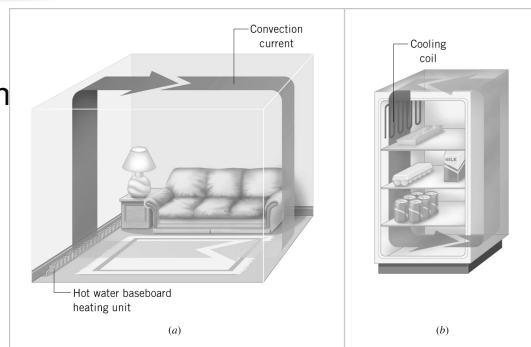
1. When a fluid is warmed:
  - Volume expands (thermal volume expansion)
  - Density decreases
2. Buoyant forces exerted by cooler (denser) fluids causes warmer fluids to rise
  - Remember Archimedes' Principle ???
3. As warmer fluids rise they cool and descend warmer fluids beneath push them out of the way
4. The net result is a natural mixing that occurs, called convection
  - Convection is a very efficient form of thermal transfer
  - Heat energy gets rapidly dispersed throughout the bulk of a fluid
5. When mixing is induced artificially (i.e. with a fan) convection occurs more rapidly & efficiently, this is called forced convection

## Examples of Convection



1. Convection currents in a saucepan

2. Convection Currents in heating & cooling appliances



## Radiation

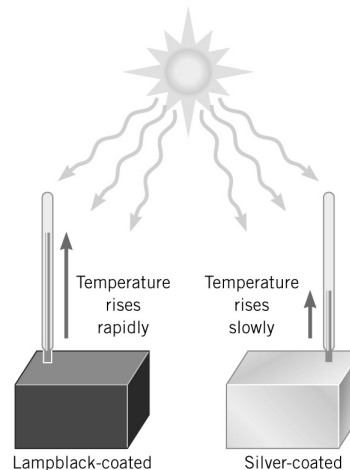
1. All objects emit(or radiate) energy in the form of electromagnetic waves
2. The rate of radiant energy emitted is related:
  - a. Surface area (A)
  - b. The 4<sup>th</sup> power of the surface temperature (T<sup>4</sup>) **{in Kelvin!!}**
  - c. The ability of an object absorb/emit radiant energy, called emissivity (ε)
    - ε = 0 for a perfect reflector
    - ε = 1 for a absorber/emitter
    - 0 < ε < 1 for normal substances and varies with wavelength
3. Combined together we have Stefan-Boltzmann's Law of Radiation:

$$\frac{dQ}{dt} = \epsilon \sigma A T^4 \quad \{\text{Rate of Radiant Heat Flow in W}\}$$

where  $\sigma = 5.67 \times 10^{-8} \text{ W} \cdot \text{m}^2 \cdot \text{K}^4$  (the Stefan-Boltzmann Constant)

## Emissivity and Absorption of Radiant Energy

1. All bodies emit as well as absorb radiant energy
2. When a body is in thermal equilibrium:
 
$$Q_{\text{absorbed}} = Q_{\text{emitted}}$$
3. Dark objects will reach higher equilibrium T than lighter objects (e.g. consider pavement and concrete on a hot day!)
  - a. Good emitters are also good absorbers of radiant energy (ε ~ 1)
  - b. Poor emitters are also good poor of radiant energy (ε ~ 0)



## The Greenhouse Effect

1. The presence of our atmosphere helps the Earth maintain a moderate & livable temperature range, due to a process called the Greenhouse Effect:
  - a. Radiant energy (short wavelength) from the sun is absorbed by the surface of earth
  - b. The earth's surface radiates energy (long wavelength)
  - c. Some of the long wavelength radiation is reflected by the earth's atmosphere by so called "Greenhouse Gases"
  - d. The reflected long wavelength radiation is reabsorbed by the earth's surface
2. This positive feedback process heats the surface of the earth and keeps the surface warmer at night
3. Do not confuse Greenhouse Effect (Good!) with Global Warming (potentially Bad!)
  - a. Global warming is believed to be due to an increased production of greenhouse gases which may increase the amount of long wavelength radiation reflected back to the Earth from the atmosphere(