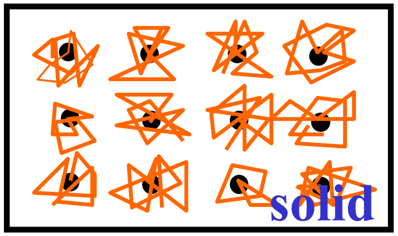
Heat (Chapter 10)

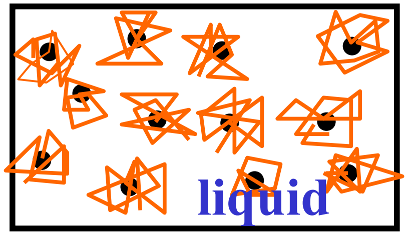
- What is heat?
- What is the relationship between quantity of heat and temperature?
- What happens to a body (solid, liquid, gas) when thermal energy is added or removed?

Thermal Energy

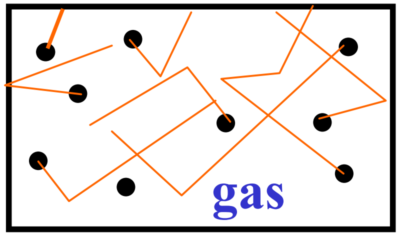
Solid: Atoms vibrating in all directions about their fixed equilibrium (lattice) positions. Atoms constantly colliding with each other.



Liquid: Atoms still oscillating and colliding with each other but they are free to move so that the long range order (shape) of body is lost.



Gas: No equilibrium position, no oscillations, atoms are free and move in perpetual high-speed "zig-zag" dance punctuated by collisions.



Heat

- Heat is a manifestation of **atomic motion** (in form of **kinetic energy** of atoms).

e.g. When rub hands together we agitate atoms at surface (by friction) causing their K.E. to increase and hence to raise temperature.

- However, in mid 1700's it was postulated that heat flow was due to an invisible, indestructible fluid called **caloric**.
- We still say "*pour on the heat*" or "*soak up the heat*" as if it were indeed a liquid!
- It was thought that the amount of caloric in an object was indestructible and uncreatable.
- However, in a clever experiment Count Rumford (Benjamin Thomson) showed that heat could be generated indefinitely (when boring a cannon barrel) which would require an infinite supply or caloric.
- Death blow to caloric theory!

Temperature and Heat

- **Temperature** of a body is a measure of the **average kinetic energy** of its constituent atoms.
- **Temperature** of a substance is **not** dependent on the **type** and **number of atoms** present (i.e. mass).

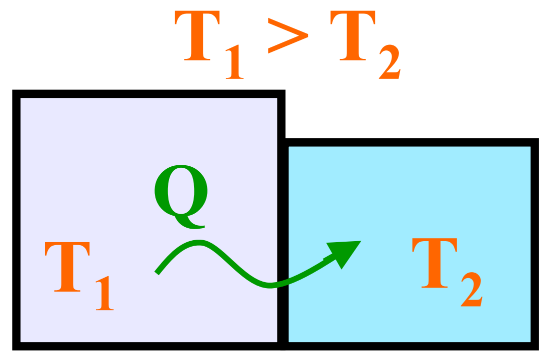
In contrast:

- **Heat** (thermal energy) of a body is a measure of the **total net K.E.** possessed by the **constituent atoms /molecules** in motion.
- The thermal energy depends on the total number of atoms present and on their average K.E. (i.e. temperature).

Example 1: The thermal energy (heat) stored in the Atlantic Ocean is enormous but the average K.E. of its molecules is low and hence its temperature is low.

Example 2: In nuclear reactions the temperature of a gas can be millions of degrees but its mass is very low and hence its heat content is low.

- Temperature also indicates whether or not and in which direction heat will flow.
- If objects at same temperature they are in thermal equilibrium and no net heat flow.



Quantity of Heat (Q):

- Heat is **energy** that flows from one object to another when there is a **temperature difference**.

Calorie (Unit of Heat):

❖ **A calorie is the amount of heat that must be added to raise the temperature of 1 gram of water by 1 ºC.**

i.e. A calorie of heat will raise water temperature by 1 ºC per gram, regardless of its temperature (range 0 – 100 ºC).

$$\boxed{Q = c.m.\Delta t} \quad \text{or} \quad Q = c.m.(T_f - T_i)$$

where: ‘c’ is a constant called “Specific Heat” of substance.

Example:

If 1000 calories of heat is added to a glass of water (270 grams) initially at 20 ºC, what is its resultant temperature?

By definition: $c = 1$ calorie /gram .ºC

$$Q = c.m.\Delta T$$

$Q = 1000$ calories
$m = 270$ grams
$T_i = 20$ ºC

$$\text{or } \Delta T = \frac{Q}{c.m} = \frac{1000}{1 \times 270}$$

$$= 3.7\ ^{\circ}C$$

$$T_f = (T_i + \Delta T) = 23.7\ ^{\circ}C$$

Note: In "metric" units 1000 calories would be required to raise temperature of 1kg of water by 1 ºC.

Specific Heat Capacity

- As **different substances** are made of **different atoms**, there is no reason to expect that the **amount of energy** required to raise temperature of given mass (e.g. 1 kg) by 1 ºC **will be the same.**
- Measurements have shown that each substance changes temperature by a characteristic amount for given heat input.

Example:

- Less heat is required to change the temperature of 1 kg of steel by 1 ºC than the same mass of water!
- Steel therefore has a lower specific heat capacity.

❖ **Specific heat capacity is a property of the material and is the quantity of heat needed to change a unit mass by one (1) unit of temperature.**

Units: **cal /gram.C** or **kcal /kg.K** or **kJ / kg.K**

Examples of Different Specific Heat Capacities:

	Substance	kJ / kg.K	kcal / kg.K	cal / g.C
High	* water	4.186	1.0	1.0
	Ice	2.1	0.5	0.5
	Steam	2.1	0.5	0.5
	Glass	0.84	0.2	0.2
Low	* Mercury	0.14	0.0033	0.0033
	Steel	0.5	0.11	0.11
	Aluminum	0.9	0.21	0.21
High	*Hydrogen	14.2	3.39	3.39
	Wood	1.8	0.42	0.42

Water (and hydrogen gas) have very high specific heats.

Practical Consequences:

- Much easier to heat up 1 kg of steel to 100 ºC than to heat 1 kg of water to boiling point!
- Water contains much more heat (about 10 times) than steel for same temperature change.
- The large specific heat capacity of water acts as **climate moderator** (maritime climate).
- Any large body of water will require a lot of heat to change its temperature (and hence temperature of surrounding air).

Result: Nights warmer and days cooler by the sea or a large lake than inland locations.

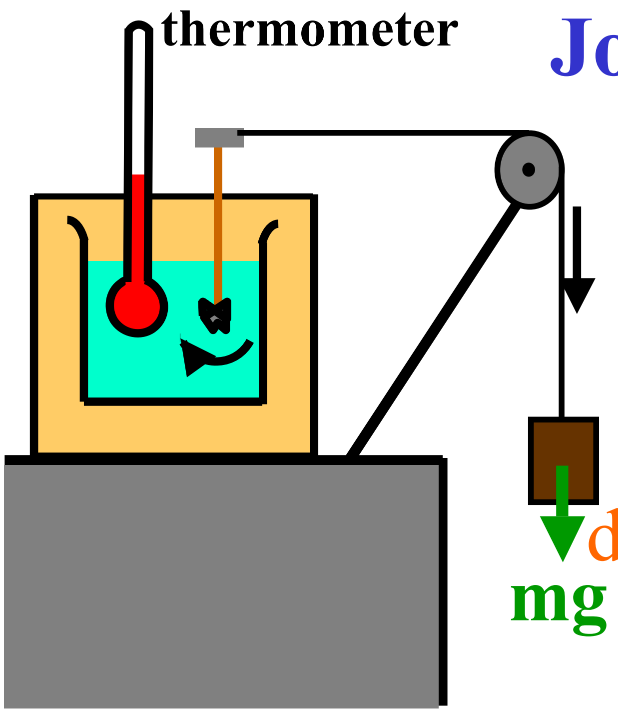
Mechanical Equivalent of Heat

i.e. The numerical relationship between **work** and **heat**.

Question: Can the temperature of a body be increased without the flow of heat from another (warmer) body in contact?

Benjamin Thompson (Count Rumford) U.S.A.

- Experimented while drilling cannon barrels for the king of Bavaria. (1798)
- Showed he could make **water boil** simply by placing it in **contact** with gun barrel.
- So heat was flowing into water from somewhere!
- 40 years later **James Joule** (UK) performed experiment **quantifying** how **heat** results from **mechanical work**.



Joule's Experiment:

- Beaker of water insulated in a box.

- Falling weight:

- Weight causes paddle to rotate doing work on liquid.

• Joule carefully measured the **temperature increase** for given **work input** (which equals P.E. change = m.g.h).

❖ **Result: 4.19 J of work raised 1 gram water by 1 ^{o}C.**

This implies 4.19 J of work is equivalent to 1 calorie of heat.

❖ **Conclusion:** As the thermal state of the system was the **same** whether the temperature was raised by **heat** of by doing **mechanical work**, this indicated their **equivalence**.

First Law of Thermodynamics

- If energy is added to a system either as **work** or as **heat**, the **internal energy** is equal to the net amount of heat and work transferred.
- This could be manifest as an increase in **temperature** or as a "**change of state**" of the body.
- First Law definition:

❖ The increase in internal energy of a system is equal to the amount of heat added **minus** the work done **BY** the system.

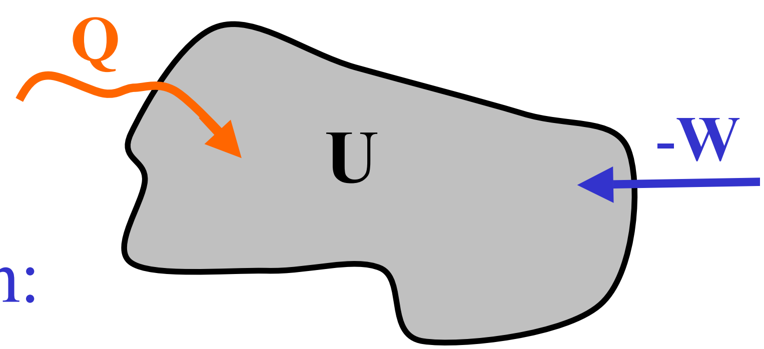
$$\boxed{\Delta U = Q - W}$$

ΔU = increase in internal energy
Q = heat
W = work done

Note: Work done **on** a system **increases** ΔU**.**
Work done **by** system **decreases** ΔU.

- The first law of thermodynamics is really a statement of **conservation of energy.**
- **Heat flow** is a transfer of **kinetic energy**.
- Either **heat** or **work** can change the **internal energy** of a system.

Internal Energy (U):



- The internal energy determines the state of a system.
- Internal energy increase can result in:
  - **Increase of temperature**
  - **Change of phase** (solid -> liquid , liquid -> gas)
- During a temperature increase K.E. increases.
- During a phase change P.E. changes (no temperature change).
- Internal energy is therefore the sum of kinetic and potential energies of atoms comprising system.
- During a **phase change** atoms are pulled further apart raising their average potential energy but **no temperature change!**