2–13 Thermal Physics

Temperature: Average KE of the particles. Depends on average **speed only** of particles (atoms or molecules)
A bucket of water at 50° has the **same temperature** as a cup of water at 50°

Thermal Energy: Average KE and the mass of the particles Depends on **speed and the mass** of the particles.

A bucket at 50° has **more thermal energy** than a cup at 100°. Bucket of $50 bills has more money than a cup of $100's. While the particles are going faster in the cup, there are so many more in the bucket.

Heat: Transfer of energy between objects that have different temperatures. The direction of heat flow depends on temperature. From hot objects (more energy) to cold objects (less energy).

Thermal Expansion: As objects are heated the molecules move faster and they expand. This is why the liquid in a thermometer rises when heated. It is how a thermostat works. It is why there are gaps in sidewalks, bridges, & railroad tracks. It is why they don't fill bottles to the top, they might explode if the liquid expands to much. The change in length is $\boxed{\Delta\ell = \ell_0 \alpha \Delta T}$. Original length ℓ_0 is multiplied by α, the coefficient of linear expansion, and the change in temp.

How long will a 5 m section of steel expand ($\alpha = 12\times10^{-6}$, from table 13-1, p 343) to if it is heated from 15° to 25°C?

$\Delta\ell = \ell_0 \alpha \Delta T = (5)(12\times10^{-6})(25-15) = 0.0006m$ so it will be $5m + 0.0006m = 5.0006m$

Kinetic Molecular Theory: Gas molecules collide with each other and their surroundings.

1. Large number of gas molecules (N) moving in random directions and a variety of speeds.
2. They are far apart, with the separation distances being vast compared to the diameter of each particle.
3. Molecules obey laws of mechanics. They attract each other, but we ignore this since the speeds and KE are huge.
4. Collisions with each other and container walls are assumed to be perfectly elastic. Conservation of p and KE.

The higher the temperature the faster the molecules move. The particles vary in speed, so we can only measure an average. Half will be going faster and half slower than the average.

Average KE reflects this speed. $\boxed{\overline{KE} = \frac{3}{2}k_B T}$ where the Boltzman Constant is k_B = 1.38x10^{-23}. Average KE is directly proportional to temperature. But velocity is not. Since $\overline{KE} = \frac{1}{2}mv^2$ and $\overline{KE} = \frac{3}{2}k_B T$,then $\frac{1}{2}mv^2 = \frac{3}{2}k_B T$ holds true. Simply and rearrange for velocity $v = \sqrt{\frac{3k_B T}{m}}$. So we can calculate the speed of the molecules on average $\boxed{v_{rms} = \sqrt{\frac{3k_B T}{\mu}} = \sqrt{\frac{3RT}{M}}}$.

This is known as the ***Root Mean Square Velocity***. μ is the mass of a single molecule and is used with the Boltzman Constant to determine the average velocity of a single molecule. M is molar mass and is used with the gas constant to find the average speed of a large group of molecules.

Pressure: $\boxed{p = F/area}$ Measured in N/m^2 called a **Pascal**. Collisions against an object are felt as pressure.

Universal Gas Law: $\boxed{pV = nRT}$ (Static problems) or $\boxed{\frac{p_0 V_0}{n_0 T_0} = \frac{pV}{nT}}$ (Changing problems)

Pressure: $\boxed{1atm = 1\times10^5 Pa}$. Volume is measured in meters cubed $1000L = 1m^3$. Gas constant $\boxed{8.31\frac{J}{mol\cdot K}}$

Static Problems: One container under a specific set of conditions. What is the temperature of 30 mol of gas at 4 atm of pressure occupying 200 L? $PV = nRT$ substituting values $(4\times10^5 Pa)(0.200m^3) = (30mol)(8.315\,J/mol\cdot K)T$

Changing Problems: One container under changing conditions, or gas moved from one container to another.

Boyle's Law: Pressure and Volume are inversely proportional. As one goes up the other goes down. 200 L of gas at 4 atm of pressure is moved to a 400 *L* container, what is its new pressure? (*n* and *T* aren't mentioned, so they must stay the same, and must cancel). $\frac{p_0 V_0}{n_0 T_0} = \frac{pV}{nT}$ simplifies to $\boxed{P_0 V_0 = PV}$. Substitute values $(4\times10^5 Pa)(0.200m^3) = P(0.400m^3)$

Charles' Law: Temperature and Volume are directly proportional. As one goes up the other goes up. 200 L of gas at 273 K are heated to 373 K, what is the new volume? (P and n aren't mentioned, so they must stay the same, and must cancel).
$\frac{p_0 V_0}{n_0 T_0} = \frac{pV}{nT}$ simplifies to $\boxed{\frac{V_0}{T_0} = \frac{V}{T}}$. Substituting values $\frac{0.200m^3}{273K} = \frac{V}{373K}$

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Combined Law: 200 L of gas at 3 atm and 273 K are moved to a 400 L container and heated to 373 K. What is the pressure? (n isn't mentioned, so they must stay the same, and must cancel)

$\frac{p_0V_0}{n_0T_0} = \frac{pV}{nT}$ simplifies to $\boxed{\frac{p_0V_0}{T_0} = \frac{pV}{T}}$. Substituting values $\frac{(3\times10^5 Pa)(0.200m^3)}{(273K)} = \frac{P(0.400m^3)}{(373K)}$

Heating and cooling curve

Specific Heat Capacity: c, Ability to absorb or retain heat. $\boxed{Q = mc\Delta T}$ where Q is heat.

Latent Heat: L, energy required to break intermolecular forces causing a phase / state change. $\boxed{Q = mL}$

Latent Heat of Fusion: L_F, energy to melt a solid. ***Latent Heat of Vaporization:*** L_V, energy to turn a liquid into a gas.

<u>When phase / state is changing temperature cannot rise since the energy is required for the change.</u>

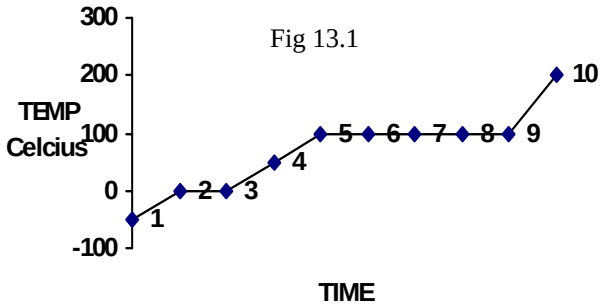

Fig 13.1

1-2	Solid	$Q = c_{solid}m\Delta T$
2-3	Melting/Freezing	$Q = mL_{fusion}$
3-5	Liquid	$Q = c_{liquid}m\Delta T$
5-9	Vaporizing/Condensing	$Q = mL_{vaporization}$
9-10	Gas	$Q = c_{gas}m\Delta T$

If more than one equation is used add the energies obtained to get the total energy.

How much energy does it take to melt 200 g of ice at 0^0C and heat it to 60^oC? It has to start as a solid, so it starts at point 2 on the graph. Melting is fusion so L is the Heat of Fusion for water from table 14-2, p. 377. Then it rises in temperature, so it must end around point 4. It is a liquid so you need the specific heat, ***c***, of liquid water in table 14-1, p. 373. (Every substance has different melting plateau, different vaporizing plateau, and different specific heats. Also each substance has different specific heats for the three principle states of matter. This values must be given in chart form or be stated in the problem.)

Melting: $Q = mL_{fusion}$ $\quad Q = (0.200kg)(3.33\times10^5\ J/kg) = 66600J$

Heat: $Q = c_{liquid}m\Delta T$ $\quad Q = (4180\ J/kg\cdot^o C)(0.200kg)(60^oC - 0^oC) = 50160J$

Total is $66600 + 50160 = 116760J$

Thermodynamics

System and environment: Think of the amount of energy in universe as constant. But, energy can be added to or subtracted from a system (engine, etc.). Energy comes from or goes to the greater universe (environment). Energy can transfer between systems. A ball colliding with another can transfer energy to that ball. Energy can change forms. If you calculate potential energy that energy can change into kinetic energy, electrical energy, thermal energy, etc.

Internal Energy: All matter has some amount of energy. Objects feel warm to the touch (internal energy: due to heat content or thermal energy of the object) caused by vibration of atoms that make up the object. If energy is added to an object it can start moving or become hotter. If energy is subtracted it can slow down or become colder. Or it can be a combination of motion and heat. Most of the beginning problems we do are done pretending that no energy is lost to the environment in the form of friction and air resistance. We often assume the energy stays in one form or that 100% is transferred when energy changes form. In real life energy is lost to the environment. Also objects become a little hotter in collisions increasing internal energy. 100% of the energy put into problems will not be available to cause motion, electricity, or power a heat engine. So many of the problems are not realistic. They ask for a theoretical unattainable maximum assuming no energy loss or perfect energy transfer.

Heat Transfer: Heat transfer occurs in one of three ways. **Conduction**: Objects touch transferring energy.
Convection: Fluids (gases & liquids) carry heat as they flow. Hot air or water rises. Convection currents.
Radiation: Electromagnetic radiation, as in infrared frequencies of light. Why does a thermos have an inner container separated from the outer by a vacuum? Why is the inside reflective?

First Law of Thermodynamics: Statement of conservation of energy. A ball falls turning PE into KE. It hits the ground causing the molecules to vibrate. The KE is transferred to internal energy and the ball comes to a stop. Have you ever seen the reverse? If internal energy caused the ball to rise, energy would still be conserved, so why isn't it reversible.

$\boxed{\Delta U = Q + W}$. ΔU : change in the system's internal energy. Q : heat added. W : work done on system (i.e. added).

+Q: heat added to system. –Q: heat lost by system. +W: work done on system. –W: means work performed by system.

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There are four basic processes that affect the outcome of the first law. These can also be visualized as *pV* diagrams.

Adiabatic: No heat enters or leaves system, then system is said to be perfectly isolated from its environment. The process happens so rapidly so that heat does not have time to flow in or out of the system. If a gas filled cylinder is compressed very rapidly, heat won't have time to escape. If $Q = 0$ is substituted into the first law $\Delta U = 0 + W$ then $W = \Delta U$.

Isothermal $\Delta T = 0$**:** Temperature of a system does not change during a process. ***The process proceeds so slowly that temperature rise is negligible***. Compress a gas filled cylinder so slowly that molecules don't speed up and temperature will not rise.

Remember, internal energy is directly affected by temperature $\overline{KE} = 3/2\, k_B T$.

If $\Delta T = 0$, then $\Delta U = 0$.

If this is substituted into the first law $0 = Q + W$, then $W = -Q$.

When T is constant Combined Gas Law reduces to Boyle's Law

$$\frac{p_0 V_0}{T_0} = \frac{pV}{T} \qquad p_0 V_0 = pV$$

It graphs as a hyperbolic function. Every T has its own isothermal curve. Higher temperature curves are farther from origin.

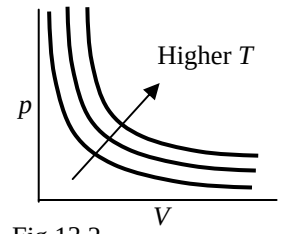

Fig 13.2

Isobaric $\Delta p = 0$**:** Pressure of a system does not change during a process.

Pressure $p = F/Area$

Multiply by a change in volume

$$\frac{F}{Area\left(m^2\right)} \cdot \Delta V\left(m^3\right) = F \cdot d\left(m\right) = W$$

So $W = -p\Delta V$

$W = -p\Delta V$

$W = -p\left(V - V_0\right)$ which implies that ***Work is the area under the pV diagram, just as Work is the area under the Fd curve.***

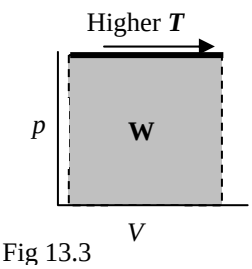

Fig 13.3

Isovolumetric (Isochoric) $\Delta V = 0$**:** Volume of a system does not change during a process.

The piston does not move so there is no force through a distance. No work.

If $\Delta V = 0$ is substituted into $W = -p\Delta V$

then $W = 0$

There is no area under the curve shown to the right.

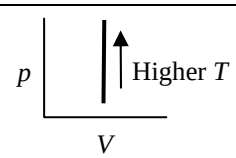

Fig 13.4

pV Diagrams and Heat Engine Cycles

***T* at any point:** Use Ideal Gas Law to find temperature at any point

$pV = nRT$

Process AB is Isobaric: $W = -p\Delta V$ and $Q = nc_p \Delta T$

Process BC is Isovolumetric: $W = 0$ and $Q = nc_V \Delta T$ and

$\Delta U = nc_V \Delta T$

Process CA isothermal: $W = -Q$

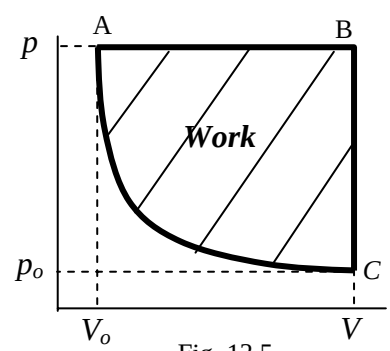

Fig. 13.5

Entire Process ABCA: In the entire process you return to point **A**. The significance of this is that you are back to the starting temperature. This means $\Delta T = 0$ for the process, so $W = -Q$ for the entire process. ***Graphically the work for an entire process is the area bounded by all the curves.***

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Carnot Engine *pV* Diagram

Hypothetical engine that runs equally well forward and backward. It operated between temperatures (isothermal lines). During the adiabatic phases it exchanged heat at hot temperature Q_{hot} and at cold temperature Q_{cold}.

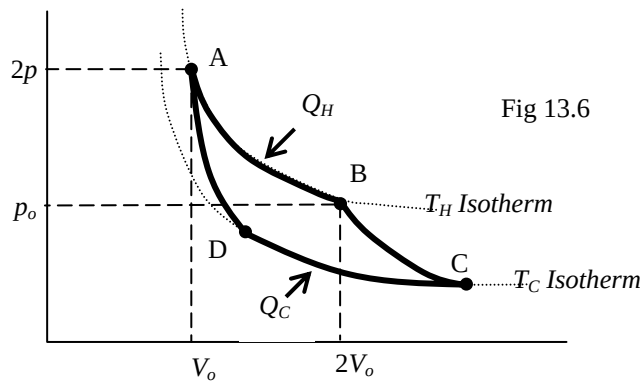


Fig 13.6

***T* at any point:** $\boxed{pV = nRT}$

Process AB isothermal: $\boxed{W = -Q_H}$

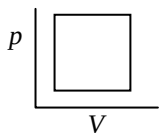
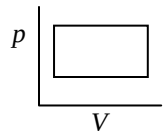
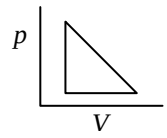
Process BC adiabatic: $\boxed{W = \Delta U}$

Process CD isothermal: $\boxed{W = -Q_C}$

Process DA adiabatic: $\boxed{W = \Delta U}$

Entire Process ABCDA: Return to point A and the starting temperature. This means $\boxed{\Delta T = 0}$ for the process, so $\boxed{W = -Q}$ for the entire process. ***Work of entire process is the area bounded by all the curves.***

Other *pV* Curves

			Any points that fit the equation $\boxed{p_0V_0 = pV}$ are on the same isothermal line. All horizontal lines are isobaric $\boxed{W = -p\Delta V}$. All vertical lines are isovolumetric $\boxed{W = 0}$. The work of a single process is the area under the curve. The work of an entire process in the area bounded by the curve.

Heat Engines: Mechanical energy can be obtained from thermal energy, only when heat is allowed to flow from a high temperature to a low temperature. Example: steam engines, internal combustion engines, and human respiration. A piece of cold wood has high thermal energy (fuel) that can be burned. This is true for gasoline and the food you eat. If this energy can be tapped it can do work. There is always a by-product: the hot exhaust of a car and human body heat are the by-products or exhaust.

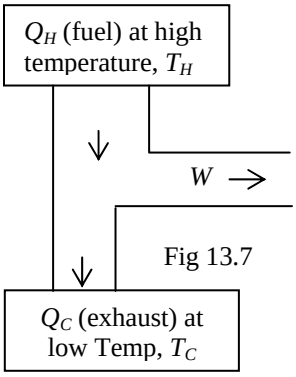


Fig 13.7

Efficiency: ratio of useful work done to the heat input.

$$\boxed{e = \left|\frac{W}{Q_H}\right|} = \frac{|Q_H| - |Q_C|}{|Q_H|} \qquad = 1 - \left|\frac{Q_C}{Q_H}\right|$$

Carnot Efficiency: an unattainable maximum theoretical efficiency. Carnot imagined a gas cylinder moved from hot to cold and expanding and contracting. He also imagined that all the processes could be done **reversibly**. Real processes involve turbulence thus making reversal along the same path impossible. Even under Carnot's ideal circumstances 100% efficiency is impossible.

$$\boxed{e_c = \frac{T_H - T_C}{T_H}} \qquad = 1 - \frac{T_C}{T_H}$$

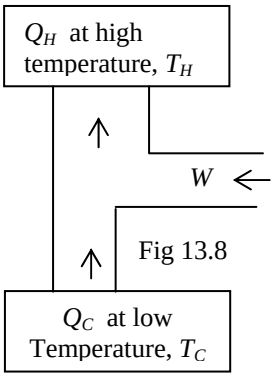


Fig 13.8

Second Law Revisited: No device is possible whose sole effect is to transform an amount of heat completely into work.

Refrigerator: (Heat Pump diagrammed to the right) Pumps heat from one area to another. Work needs to be added. But, what's happening to Q_H? It's getting bigger. So overall a refrigerators actually produce heat. As an example, in the process of lowering the inside air volume 40 degrees they may raise an equal volume of outside air by 100 degrees.

Entropy: A measure of how much energy or heat is unavailable for conversion to work. Heat is sometimes called the graveyard of energy. As Carnot showed, not all the heat energy can be used to do work. Entropy is thought of as disorganization. As we use more energy we get more heat as a by-product. Look at oil. It is composed of large well organized molecules. When we burn it to produce work, we get smaller less organized molecules and heat. The small gas by-products of combustion, CO_2 and H_2O, are more difficult to burn and extract further work from. Gases are less organized than liquids, which are less organized than solids. Perhaps all highly organized molecules will turn into small atoms of gas evenly distributed throughout the universe and they may eventually reach an even average temperature. If there is no heat difference there can be no work. So life and machines can't exist. This is called the heat death of the universe.

Second Law Re-Revisited: Natural processes to move toward a state of greater disorder (entropy). The entropy of an isolated system never decreases. Entropy can only really stay the same for idealized (reversible) processes. So it always increases. The total entropy of any system and the environment increases as a result of any natural process.

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