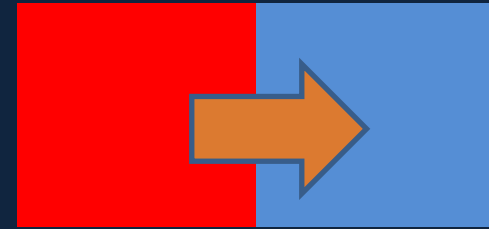


Heat and Energy Conservation

Physics 1425 Lecture 33

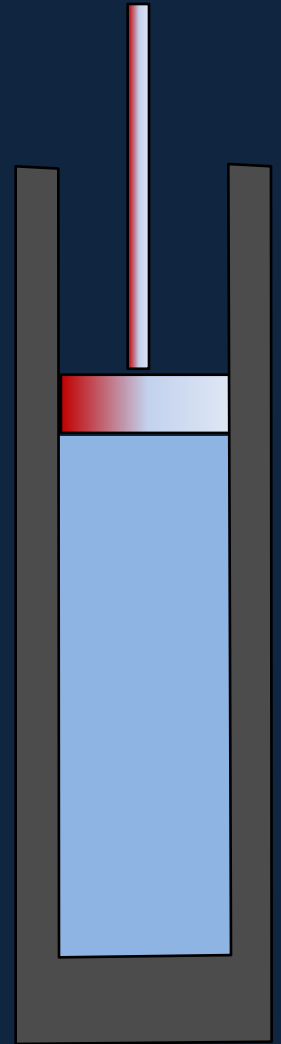
Heat Flow



- If something warm is in contact with something cooler, the warm thing cools down as the cool thing gets a little warmer. This flow of energy, called **heat**, was until the 1800's envisioned as an invisible fluid, called caloric.
- This wasn't such a crazy idea—using fluid flow equations, Fourier correctly analyzed heat flow in the Earth, inventing Fourier series to do it. And, there existed at least one invisible fluid: electricity.
- The idea was that temperature acted like a potential for the caloric fluid, and it flowed downhill to lower T .

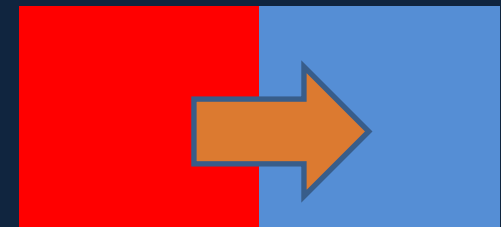
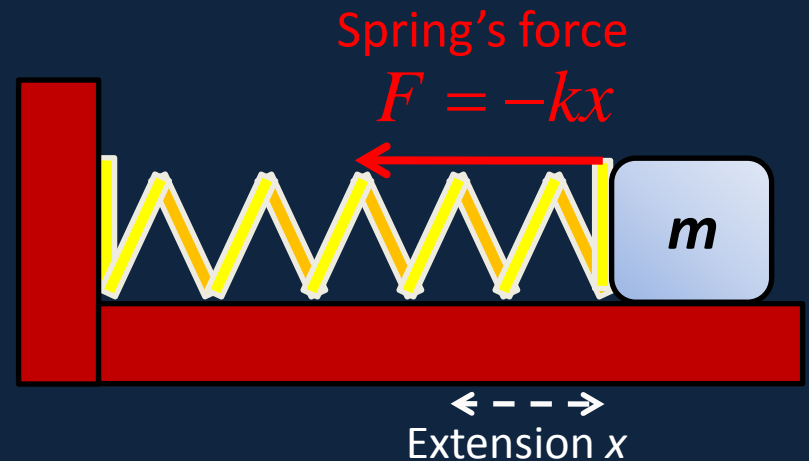
What's Wrong with a *Fluid* Heat Theory?

- If you believe that things only rise in temperature when something fluid flows in, it's hard to account for the sudden rise in temperature of gas in an insulated (no heat flow possible) container when it is suddenly compressed by pushing down the piston.
- Basically, **the caloric theory does fine in describing heat flow, even in complicated shapes, but breaks down when work is done.**



The Right Approach: Energy Transfer!

- We've discussed **work as energy transfer**: the spring does work on the mass, the spring loses potential energy as the mass gains that amount of kinetic energy.
- **Heat flow is another form of energy transfer**—we've seen that a hotter gas has more kinetic energy per molecule.



Units for Heat

- Since heat flowing from a hot body to a cooler one is just a transfer of energy, the natural unit is the same as for energy: the **joule** (J).
- The kilojoule, **kJ**, is often a convenient unit.
- An old unit, still in wide use, is the **calorie**, the heat needed to raise the temperature of 1 gram of water by 1°C . *The calories used in food labeling are actually kilocalories.*
- Finally, the **Btu**: 1 lb of water by 1°F .

Joule Heats Things Up

- James Joule, an English brewer, measured the frictional heat generated by a measured amount of work: a falling weight drove paddles under water in an insulated can.
- His result is called the **mechanical equivalent of heat:**

$$4.19\text{J} = 1 \text{ cal}$$



Internal Energy of an Ideal Gas

- We'll consider first a monatomic gas.
- Since it's ideal, we're ignoring collisions or any kind of interactions. That is, we ignore any potential energy in the system.
- The **only energy is the kinetic energy**, and we know average molecular $KE \frac{1}{2}mv^2 = \frac{3}{2}kT$, so

$$E_{\text{int}} = N \left(\frac{1}{2} m \overline{v^2} \right) = \frac{3}{2} NkT$$

- Important to notice: this **internal thermal energy per atom** **doesn't depend on the atom's mass!**

Specific Heats of Solids and Liquids

- The specific heat c (per unit mass) of a material is defined as the amount of heat energy that must be supplied to raise the temperature of 1 kg by 1°C.
- To raise the temperature of m kg by ΔT takes heat
$$Q = mc \Delta T$$
- For almost all solids and liquids, c can be taken as constant over wide temperature ranges—but it changes dramatically when the substance melts or boils, as we'll soon discuss.
- Water has a very high specific heat: 4186 J/kg·°C

Dulong and Petit Law

- In 1820, these two professors in France discovered that the product **atomic weight x specific heat was almost the same** for many very different elements.
- We can now see this means the specific heat per atom *is* approximately the same—just as it is for the ideal gas.
- Elements with **heavier atoms** have **lower specific heat per kg** *because there are fewer atoms in one kg.*

| Atom | c J/kg.°C | m | $c \times m$ |
|----------|----------------|------|--------------|
| Lead | 129 | 207 | 26700 |
| Zinc | 388 | 65.4 | 25400 |
| Sulphur | 710 | 32.1 | 22800 |
| Aluminum | 900 | 27 | 24300 |
| Silver | 232 | 108 | 25100 |
| Mercury | 140 | 201 | 28100 |

Clicker Question

- I have 100 ccs of mercury at 20°C in an insulated container. I pour in 100 ccs of water at 100°C . What is the final temperature, after thermal equilibrium is established?
 - A. 60°C
 - B. Between 20°C and 60°C
 - C. Between 60°C and 100°C

Calorimetry...

... is just thermal energy bookkeeping!

- Two or more objects at different temperatures are brought into thermal contact, and heat flows between them.
- Two objects A, B at initial temperatures T_A , T_B ($T_A > T_B$) will reach an equilibrium temperature T given by A's heat loss = B's heat gain,

$$(T_A - T)c_A m_A = (T - T_B)c_B m_B$$

Latent Heat

- A large amount of heat has to be supplied to **water** at $100\text{ }^{\circ}\text{C}$ to turn it **into steam** at $100\text{ }^{\circ}\text{C}$: **539 calories per gram**, as opposed to the mere 100 calories per gram to raise it from $0\text{ }^{\circ}\text{C}$ to $100\text{ }^{\circ}\text{C}$. *This does not change the rms speed of the molecules: it **all** goes into **potential energy**, pulling the molecules away from each other against their attractive forces.* Similarly, it takes **80 calories per gram to melt ice**. This is why ice cubes remain in a drink quite a while.
- This is **latent heat**, and must of course be **included** in calorimetry-type bookkeeping.

Evaporation

- Recall that if water is put in a closed container, initially filled with dry air, water molecules will leave the liquid until a liquid-vapor equilibrium is established, the partial vapor pressure depending on temperature.
- There is a **latent heat associated with this evaporation**, and it takes a little **more energy to escape from water at 20°C** (585 cal/gm) than at 100°C (539 cal/gm). The cooler water is denser and exerts a slightly more powerful attraction.
- **As the vapor forms, the liquid cools—it's losing its fastest molecules. This is how sweating works.**

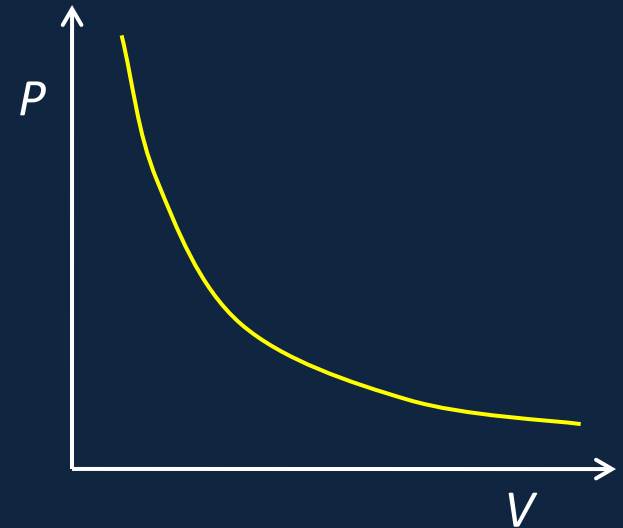
The First Law of Thermodynamics

- A closed system has a total internal energy E_{int} .
- This energy can be changed in two different ways:
 - A. The system can do **work W** , or have work done on it, $-W$.
 - B. **Heat Q** can flow into the system, or $-Q$ flow out.
- The First Law is just total energy conservation:

$$\Delta E_{\text{int}} = Q - W$$
- **Change in internal energy = heat in – work done.**

States of an Ideal Gas

- Suppose we have n moles of an ideal gas in equilibrium in a piston. The “state” of the gas can be defined by giving the **state variables P, V** . The gas state is a point in the (P, V) plane.
- If heat is exchanged or work done, the gas state variables **trace a path** in the (P, V) plane.
- If the gas moves along an **isotherm $PV = \text{constant}$** , its internal energy stays the same.

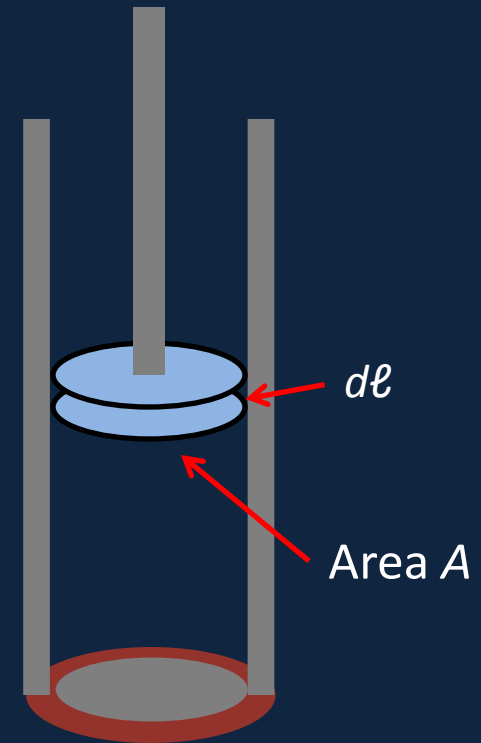


The temperature (and therefore internal energy) is constant along an isotherm, $PV = \text{const.}$

The Gas Does Some Work

- The gas pressure P means the force from the gas on the piston, of area A , is PA . This means that if the piston moves up a distance $d\ell$, the work done by the gas, force \times distance = $PA d\ell = PdV$.
- The work done by the gas in expanding **isothermally**, along $PV = nRT$, is

$$W = \int_{V_A}^{V_B} PdV = \int_{V_A}^{V_B} \frac{nRT}{V} dV = nRT \ln \frac{V_B}{V_A}$$

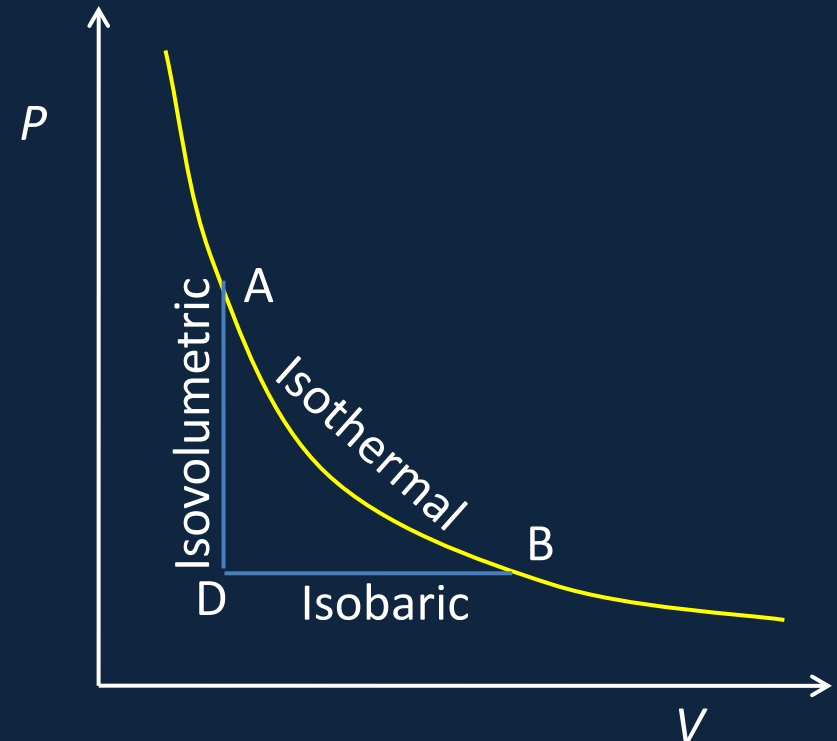


Clicker Question

- When an ideal gas is compressed isothermally, work is done. Where did that energy transfer go?
 - A. It's stored as extra pressure in the gas.
 - B. It's stored as internal energy in the gas.
 - C. It must have leaked out as heat.

Different Paths from State A to State B

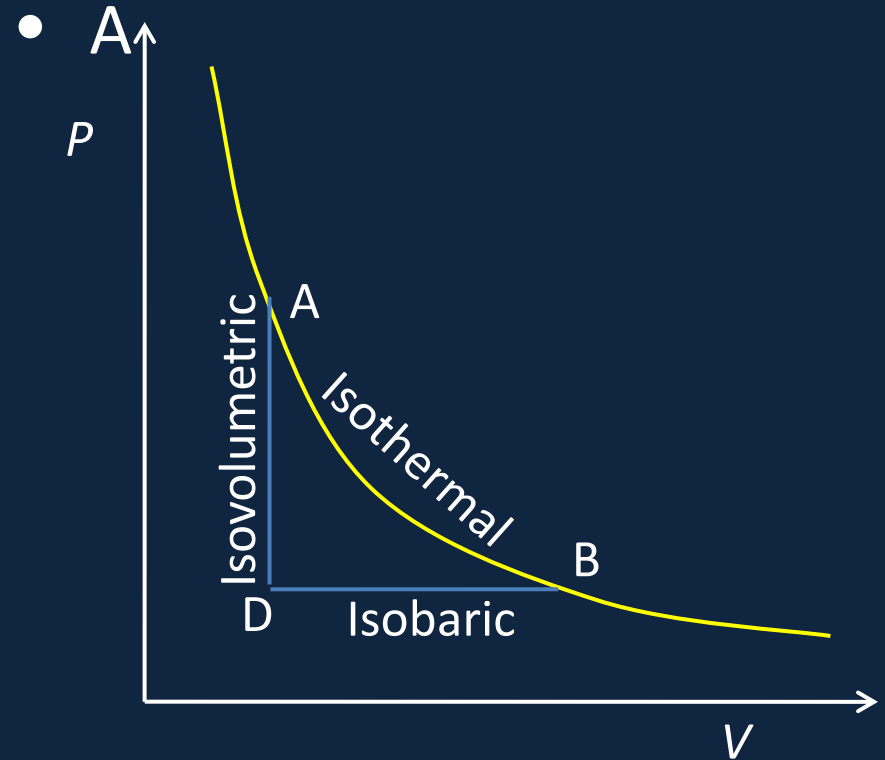
- Along the isotherm AB, internal energy is constant, $Q = W$: heat supplied = work done.
- Along the constant volume path AD *no* work is done, heat is lost.
- Along the constant pressure path DB the gas does work *and* absorbs heat.



Clicker Question

- Suppose the gas follows the path ADB. How does the internal energy lost along AD compare with that gained along DB?

- A. It's less
- B. It's the same
- C. It's more



Clicker Question

- Suppose the gas follows the path ADB. How does the total work done by the gas compare with that if it followed the isotherm from A to B?

- A. It's less
- B. It's the same
- C. It's more

