Physics 111
Lecture 24
Thursday, December 2, 2004

- Ch 17: Latent Heat
  - Ideal Gas Law
  - Kinetic Theory
  - Stress
  - Strain

**Announcements**

**The Physics 111 Help Session**

For the rest of the semester

- Sunday, Dec. 5, 6:30 - 8 pm in CCLIR 468
- Thursday, Dec. 9, 3 - 5pm and 7 - 9pm

**Announcements**

**Exam #3**

Monday, Dec. 13, 2004
10:30 am – 12:30 pm
everything

- **Hint**: Be able to *do the homework* (graded AND recommended) and you’ll do fine on the exam!
- You may bring one 8.5”X11” crib sheet (handwritten on both sides), a pencil or pen, and a scientific calculator with you.
- I will put any constants and mathematical formulas that you might need on a single page attached to the back of the exam.

**Announcements**

**Ch 17: Phases & Phase Changes**

- The review of your chemistry class continues. In this chapter, we look at phase transitions, the kinetic theory of gases, and solid deformations.

### Chapter 17

**Phases & Phase Changes**

When a substance changes *phase*, extra energy is either released or required.

We classify matter as existing in one of 3 phases:

- **solid**
- **liquid**
- **gas**
If we start out with our substance in a gaseous phase, then how do we get it to become a liquid?

Think of the result of a lower temperature in terms of the molecular argument about the speed of the molecules...

Molecules are more likely to stick together.

Energy is released as a gas becomes a liquid!

In order to get a substance to change phase from a liquid to a gas, we must...

We have to add energy to vaporize a liquid.

The amount of energy we must add to get a substance to change from a liquid to a gas is the same as the amount of energy that is released when the same substance changes phase from a gas to a liquid...and is known as the latent heat of vaporization $L_v$

On what property of our substance could the latent heat (that is, the amount of heat necessary to melt our substance from a solid to a liquid of to vaporize our liquid to a gas) possibly depend?

How much of the substance we’ve got.

Clearly, the more of it I’ve got, the more heat I’m going to have to transfer to that substance to get a given temperature change.

We specify this quantity with mass $(m)$.

Should this graph pass through the origin?

That is, how much heat must be transferred to get the phase to change for an amount $m$ of the substance?

The slope of this line is the Latent heat ($L$) of either fusion of vaporization, depending upon the phase change we’re examining.

A 100.0 g ice cube is initially at $T_i = -0.2.0^\circ C$. The ice is placed in a sealed container over a heat source. Heat is transferred to the ice until it becomes steam at $T_f = 120^\circ C$. How much total heat was transferred to the ice?

Use the following information:

- $c_{\text{H}_2\text{O}(s)} = 2090 \text{ J/kg}^\circ\text{C}$
- $L_f = 3.33 \times 10^5 \text{ J/kg}$
- $c_{\text{H}_2\text{O}(l)} = 4190 \text{ J/kg}^\circ\text{C}$
- $L_v = 2.26 \times 10^6 \text{ J/kg}$
- $c_{\text{H}_2\text{O}(g)} = 2010 \text{ J/kg}^\circ\text{C}$

$Q = mL$
Calorimetry Problems

Tips for solving such problems:

1) Make sure your units are consistent!

2) Only use the specific heat formula for temperature changes without phase changes.

\[ Q = mc\Delta T \]

Molar Mass

A property related to the molecule itself, the atomic weight is related to the molar mass.

For example, oxygen has an atomic weight of 16. This means that 1 mole of atomic oxygen will have a mass of 16 grams.

1 mole of ANY substance contains Avagadro’s number of molecules

\[ N_A = 6.022 \times 10^{23} \text{ molecules} \]

Ideal Gases

Most gases with which we deal on a day-to-day basis behave as ideal gases.

Ideal gases obey the ideal gas law, a physical law that relates the pressure of a gas, the number of molecules in the gas, the volume of the gas, and its temperature.

The Ideal Gas Law

We can reason out this relationship fairly easily. Keeping everything else constant...

What happens to the temperature of a gas if we increase its pressure?

Pressure is proportional to temperature.

\[ P \propto T \]

What happens to the pressure of a gas if we increase the number of molecules for fixed \( V \) and \( T \)?

Pressure is proportional to the number of molecules.

\[ P \propto N \]
We can reason out this relationship fairly easily. Keeping everything else constant...

**What happens to the volume of a gas if we increase the number of molecules at fixed \( P \) and \( T \)?**

Volume is proportional to \( N \).

\[ V \propto N \]

**Ideal Gas Law:**

\[ PV \propto NT \]

Summarizing, we have:

Therefore, we can confidently state that

\[ P = k_B \frac{NT}{V} \]

where \( k_B \) is the proportionality constant, **Boltzmann’s constant**.

It is also often convenient to write the ideal gas law using the **number of moles** rather than the number of molecules. In this case, we have

\[ PV = nRT \]

where \( n \) is the number of moles of gas \( (N/N_A) \) and \( R \) is the ideal (universal) gas constant.

More typically, the ideal gas law is written

\[ PV = Nk_B T \]

So a common, but equivalent form of the ideal gas law, is:
What volume is occupied by 1 mole of gas at temperature 0°C and pressure (1 atm)?

1) 0
2) 3.8 x 10^-21 L
3) 22.4 L
4) 45 L
5) 2270 L
6) None of the above

Worksheet #2

How many molecules will you find in 1 cm³ of gas at temperature 0°C and pressure of 1 atm?

1) 4.52 x 10^-12
2) 4.46 x 10^5
3) 2.65 x 10^14
4) 2.68 x 10^16
5) 2.68 x 10^19
6) None of the above

Worksheet #3

I've got a box full of molecules. They're bouncing around in my box in all different directions, at a variety of speeds.

What happens when a molecule runs into a wall of my box? It exerts a force on that wall. Can we figure out the magnitude of this force?

Let's simplify our problem by examining a 1-D gas instead of a 3-D gas. Assume our molecules only move in the x-direction, and that all collisions they make with the walls are perfectly elastic.

What is the magnitude of the force that has been applied by the right wall to our molecule?
How much time do we have to wait between applications of this force on the same molecule by the same wall?

$$\Delta t = \frac{2m}{v_{xi}}$$

We can now plug our solution for $$\Delta t$$ into our expression for force and get

This is the force of the wall on a single molecule.

By Newton’s 3rd Law, this must also equal the force of a single molecule on the wall.

So, now let’s look at the total force on the wall, which will be the sum of the forces exerted by each molecule:

$$F_{tot} = \frac{m}{d} v_{x1}^2 + v_{x2}^2 + \cdots + v_{xN}^2$$

We note that the average square velocity of our molecules is simply

$$v_x^2 = v_{x1}^2 + v_{x2}^2 + \cdots + v_{xN}^2$$

One last step…now let’s find the pressure exerted on a wall. Remember, pressure is force divided by area.

$$P = \frac{F}{A} = \frac{F_{tot}}{d^2} \frac{mNv_x^2}{d^2}$$

So we can now rewrite our force in terms of this average velocity squared quantity:

Now, we expand our analysis to a 3-D gas. In a 3-D gas, no direction is special. This implies that the mean square velocity in the $$x$$-direction equals that in the $$y$$ and $$z$$ directions as well.

But since the average speed in each direction is the same

$$v_x^2 = v_y^2 = v_z^2$$

So we can now rewrite our force in terms of this 3-D average velocity squared:

One last step…now let’s find the pressure exerted on a wall. Remember, pressure is force divided by area.
Manipulating this just a bit, we get:

\[ P = \frac{2}{3} \frac{N}{V} \left( \frac{1}{2} m < v^2 > \right) \]

So the pressure is proportional to the number of molecules per unit volume times the kinetic energy per molecule!

**Molecular Understanding of Pressure & Temperature**

Let's rewrite our equation for pressure in a form that resembles the ideal gas law:

\[ P = \frac{2}{3} \frac{N}{V} \left( \frac{1}{2} \frac{m}{k_B} < v^2 > \right) \]

Recall, the ideal gas law says:

\[ PV = Nk_B T \]

Equating the right-hand sides of these equations we get

\[ \frac{2}{3} \frac{N}{N} \frac{m}{m} \frac{v^2}{v^2} \]

\[ T = \frac{2}{3} k_B \frac{1}{k_B} \frac{1}{k_B} < v^2 > \]

OR

\[ \frac{3}{2} k_B T = \frac{1}{2} m < v^2 > \]

We now see directly the relationship between Temperature and molecular kinetic energy.

An ideal gas has an initial temperature of 20°C. The gas is heated, and the kinetic energy is observed to increase by a factor of 2. What is the final temperature of the gas?

1) 20°C  
2) 40°C  
3) 80°C  
4) 313°C  
5) 586°C  
6) None of the above

Recall that

\[ V^2 \]

\[ v_x^2 \]

So

\[ \frac{1}{2} m < v_x^2 > = \frac{1}{2} k_B T \]

This result implies that 1/3 or the kinetic energy (on average) can be found in each direction the gas molecules are allowed to move. We call these directions in which the molecules may move "degrees of freedom"—i.e., the number of ways in which a molecule is free to move.

**Ch 17: Phases & Phase Changes**  
**Phys 111**
Molecular Understanding of Pressure & Temperature

**Equipartition Theorem**

Energy is equally divided among the available degrees of freedom for a system in thermal equilibrium.

This theorem implies that the total energy of a 3-D gas with \( N \) molecules is given by

\[
E_{\text{tot}} = N \left( \frac{1}{2} m \langle v^2 \rangle \right) = \frac{1}{2} N k_B T = \frac{3}{2} nRT
\]

**RMS speed**

The close relationship between temperature and kinetic energy of the molecules in a gas allows us to find the average speed of those molecules if we know the temperature. We call this speed the root-mean-square (rms) speed.

\[
\nu_{\text{rms}} = \sqrt{\langle v^2 \rangle} = \sqrt{\frac{3k_B T}{m}} = \sqrt{\frac{3RT}{M}}
\]

\( M \) is the mass per mole

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**How would the rms speed of the molecules in a gas change if the pressure of the gas was doubled while the volume was kept constant?**

1) increase a factor of 4
2) increase by a factor of 2
3) increase by a factor of 1.4
4) stay the same
5) decrease by a factor of 1.4
6) decrease by a factor of 2
7) decrease by a factor of 4

**Worksheet #5**