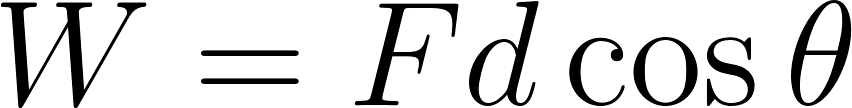
Energy “Story”

What story do we want to tell with energy?   
What does the book look like?

# What do they know from Unit 3?

* [](https://www.codecogs.com/eqnedit.php?latex=W%20%3D%20Fd%20%5Ccos%20%5Ctheta)
* Signs of work
* Work is area of F(x)
* They have heard of energy in other courses, and know that it is conserved
* Work has the units of energy

Unit 4 Introduction

# Reductionism and Emergence

*Take* [*http://umdberg.pbworks.com/w/page/68371403/Reductionism%20and%20emergence%20(2015)*](http://umdberg.pbworks.com/w/page/68371403/Reductionism%20and%20emergence%20(2015))

Chapter 12 - Introduction to the Microscopic World

# 12.0 Introduction

*Take:* [*http://umdberg.pbworks.com/w/page/68405576/The%20micro%20to%20macro%20connection%20(2013)*](http://umdberg.pbworks.com/w/page/68405576/The%20micro%20to%20macro%20connection%20(2013)) *with removing the last paragraph and adding what is below.*

In this chapter, we will look at some important properties of matter at the molecular scale such as the idea of a mole which you may know from a previous course. We will then develop molecule-based pictures of gases and solids. We will use these models of matter to help us to develop a coherent picture of energy that spans from our everyday world to the world of molecules.

# 12.1 Atomic structure and symbolism

*Take this from section 2.3 from OpenStax chemistry.*

# 12.2 The periodic table

*Take section 2.5 from OpenStax chemistry through figure 2*

# 12.3 What is a mole

*For this, let’s use section 3.1 from OpenStax Chemistry, removing only the two paragraphs at the top to make it flow*

# 12.4 Basics of Temperature

*David: Use section 13.1 from OpenStax Physics here*

# 12.5 The molecular picture of gases

## 12.5.1 The ideal gas law

*Section 13.3 removing the bit about moles (keeping the re-statement of ideal gas law in terms of moles). Also remove the bit at the end about the Ideal Gas Law and energy.*

## 12.5.2 Ideal gases vs. real gases

*Use* [*https://www.youtube.com/watch?v=3vGfWoPIstk&feature=youtu.be*](https://www.youtube.com/watch?v=3vGfWoPIstk&feature=youtu.be)

# 12.6 The molecular picture of solids

## 12.6.1 What is a solid

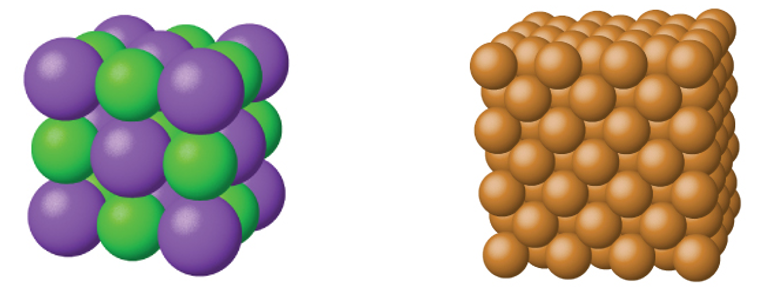
*Use* [*http://umdberg.pbworks.com/w/page/68393164/Solids%20(2013)*](http://umdberg.pbworks.com/w/page/68393164/Solids%20(2013)) *except the last paragraph*

## 12.6.2 The solid state of matter

*Use 10.5 from OpenStax chemistry here. Remove the part about defects. Also be sure to remove the discussion about lattice energies.*

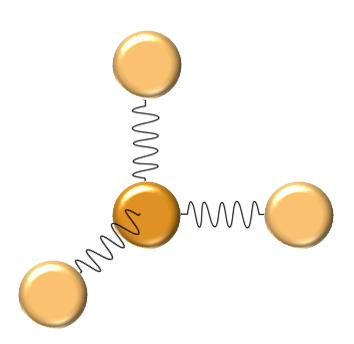
## 12.6.3 Einstein solid

In this class, we will be focused on a model called the *Einstein solid* model of simple crystalline solids, particularly those where the atoms are arranged in simple cube-based structures such as table salt (NaCl) and copper as shown in Figure **A**. This model, named after Albert Einstein, is like the ideal gas law discussed in the previous section: almost no solid behaves exactly like an Einstein solid. However, the behavior of many solids is approximately Einsteinian and the principles behind the model can be used successfully to understand the more complex solids that you will probably study in your other science courses.



*Figure* ***A****: Both table salt (NaCl) on the left and solid copper on the right are simple crystalline solids with a cube-based structure.*

In the Einstein solid model, we model the bonds between each atom as tiny springs that obey the Hooke’s Law that we discussed in Unit 2. This model is shown in Figure **B**.While chemical bonds, like everything else, do NOT exactly obey Hooke’s Law, the modelling of atomic bonds as springs is very common throughout physics, chemistry, and biology. The reason this model is so common is that, for small shifts, atoms do behave as if they are on springs: pull two atoms in a crystal apart and the atomic bonds will pull them back together. If you push two atoms too close together, then the two positively charged nuclei will repel each other pushing them apart. The result is that each atom vibrates about its equilibrium position or *lattice site* as if it were attached to tiny springs. Given that calculations with Hooke’s Law are so simple to do in comparison to full calculations of atomic forces, you can see why this Einstein model is a useful approximation.



*Figure* ***B****: The model of an Einstein solid, the central atom is connected to its neighbors by atomic bonds that are modeled by springs that obey Hooke’s Law*

Chapter 13 - Introduction to Energy

This chapter introduces a lot of terms and basic ideas. These terms and ideas will be elaborated upon in the later chapters.

# 13.1 Introduction - What *is* energy and how is it different from forces?

*Use video* <https://youtu.be/EXf5gtyKQ8Q>

# 13.2 Units of Energy

If energy is defined as the ability to do work, then energy and work must have the same units. Thus, the SI unit of the energy is the Joule (recall ). Energy, however, is one quantity where there are many other units in common use in scientific literature including electron-Volts (eV), kilowatt-hours (kW∙hr), calories, and Calories. In this course, we will be using Joules and electron-Volts exclusively. We are including these other units for your reference.

## Electron-Volts

A common quantity in chemistry is the electron-Volt or eV. One electron-Volt is the amount of energy gained by an electron as it travels between the two ends of a 1 Volt battery (a concept that will be discussed in more detail when you study electricity). Numerically, 1eV = 1.602×10-19J. The reason this unit is common in chemistry is that the energies of atomic bonds are typically about 1eV as shown in the table below[[1]](#footnote-0). The bond-dissociation energy is the energy *released* when the bond is formed.

|  |  |  |
| --- | --- | --- |
| Bond | Bond-dissociation energy at 298K (eV/Bond) | Comment |
| C-C | 3.60-3.69 | Strong, but weaker than C–H bonds |
| Cl-Cl | 2.51 | Indicated by the yellowish colour of this gas |
| H-H | 4.52 | Strong, nonpolarizable bond  Cleaved only by metals and by strong oxidants |
| O-H | 4.77 | Slightly stronger than C–H bonds |
| OH-H | 2.78 | Far weaker than C–H bonds |
| C-O | 11.16 | Far stronger than C–H bonds |
| O-CO | 5.51 | Slightly stronger than C–H bonds |
| O=O | 5.15 | Stronger than single bonds  Weaker than many other double bonds |
| N=N | 9.79 | One of the strongest bonds  Large activation energy in production of ammonia |
| H3C-H | 4.550 | One of the strongest aliphatic C–H bonds |

## Kilowatt Hours

When you buy electricity from the power company, the bill says how many kilowatt hours you have purchased. A Watt is a unit of a quantity called *power* and 1 Watt is equal to 1 Joule/second: 1W = 1 J/s. Thus, a kilowatt hour is:

## Calories and calories

The calorie is an imperial unit of energy that is still in common use in the nutritional sciences in the United States. One calorie (lowercase c) is the amount of energy needed to raise 1g of water 1oC or 1 cal = 4.814J. On food labels, you will see energy listed in Calories (capital C). One Calorie is equal to 1kilocalorie; in other words, 1 Cal = 1000 cal. Thus, one 1 Cal = 4814 J. In other countries, you will see food labels in both Calories and Joules like the one shown in Figure **A[[2]](#footnote-1)**.



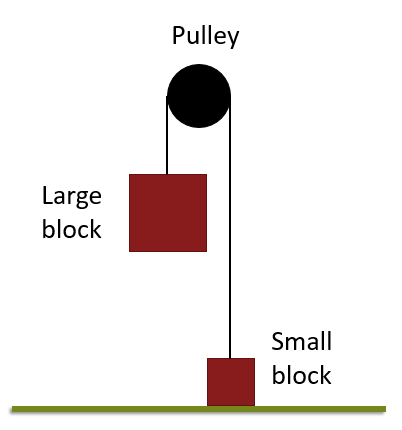
*Figure* ***A****: A food label from the UK showing the energy of the food in both Joules and kcal (or Calories).*

# 13.3 Types of Energy and Scales of Energy

Fundamentally, there are only two kinds of energy: kinetic energy and potential energy*. Kinetic energy (K) is the ability to do work associated with motion and potential energy (U) is the ability to do work arising from the relative positions of two or more objects.* As an example, the car in motion in the left image of Figure **B** has the capability to do work due to its motion - the car has *kinetic energy*. If the car were to crash, then a force would be exerted over a distance deforming the car (right image in Figure **B**). The sheer fact that the car is moving means that it *can* do work. Similarly, the larger block in Figure **C** could do work if the system were released. As the large block fell, it would lift the small block. The large block has *potential energy* - an ability to do work due to its position relative to the earth.

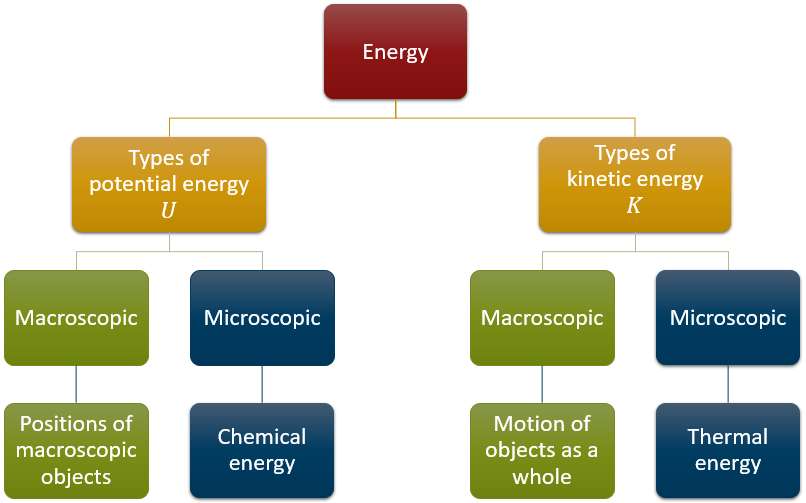


*Figure* ***B****: A car traveling down the road (left) has an ability to do work due to its motion - it has kinetic energy. We see that ability to do work when the car crashes (right) - a force acts for a distance deforming the car. [[3]](#footnote-2)*

**

*Figure* ***C****: A large block connected to a small block over a pulley has an ability to do work due to its position relative to the earth; the large block has potential energy. We see that work when the large block is released exerting, via the rope and pulley, a force on the small block for a distance causing it to accelerate upwards.*

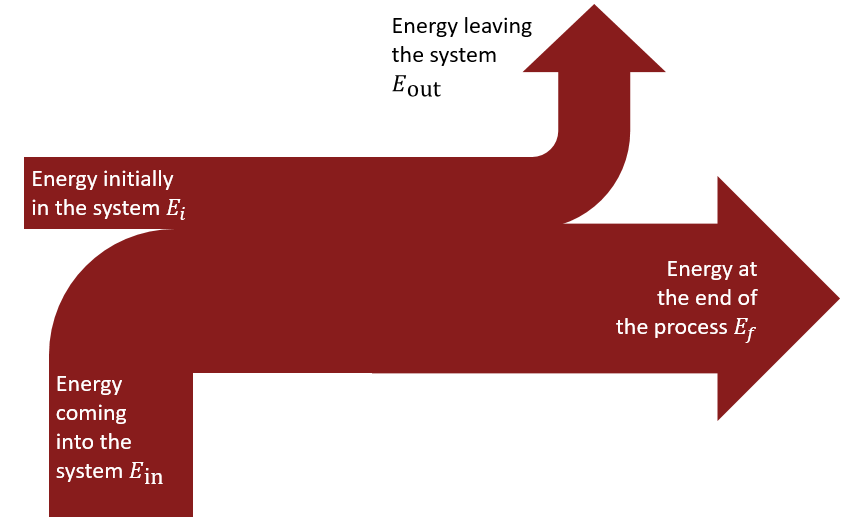
All of the many different types of energy that you have heard about in previous courses, thermal, chemical, electrical, etc., all ultimately boil down to these two different types. You may be wondering how chemical and thermal energy can be potential or kinetic. Typically when we think of kinetic energy, we think of the motion of people, cars, and the like! The key is to think about the *scale* of the energy: are we talking about energy at the macroscopic scale (people etc.) or the microscopic scale (atoms and molecules)? As we shall see, thermal energy is just kinetic energy on the microscopic scale and chemical energy is potential energy on the microscopic scale. The relationships between these types of energy can be seen in Figure **D**. One of our goals throughout these chapters on energy is to develop a coherent picture of energy that applies at both the macroscopic scale of people etc. and at the microscopic scale of atoms and molecules. Thus, while we may present the macroscopic and microscopic scales in two separate chapters, keep in mind that we are talking about the same idea of energy throughout. At the end, we will look at how to transfer energy between these two different scales.



*Figure* ***D****: The relationships between different types of energy*

# 13.4 Conservation of Energy

You probably have heard from other courses that “energy is conserved.” This statement is true, for the Universe as a whole - the total amount of energy has not changed since the birth of the Universe 13.6Gyr ago. While this is a hugely important fact which gets deep into the heart of physics, it may not seem very useful. However, there is an, equally fundamental, and more useful fact: the amount of energy in any given system is conserved: the energy at the end of some process is what I had at the beginning plus any that came in minus any that went out, or in equation form: .



*Figure* ***E****: The conservation of energy in graphical form*

# 13.5 Ways to Transfer Energy

So what ways are there to transfer energy into or out of a system? Well we already know of one way: work. If we do positive work on a system (the force we apply is in roughly the same direction as the displacement), then we will add energy *in*. Conversely, if we do negative work on a system (force essentially opposing the displacement) then energy is leaving the system. This should make some intuitive sense: we expect for positive work that the object will speed up - the object will gain kinetic energy. On the other hand, if we do negative work, we expect that the object will lose kinetic energy.

There is another way to transfer energy into or out of a system: heat which we represent by the letter .

*Use* [*https://youtu.be/95AUDZn9bZo*](https://youtu.be/95AUDZn9bZo)

In summary, heat is the transfer of energy by collisions at the microscopic scale. Heat has nothing to do with temperature*.* Since heat is a transfer of energy into or out of a system, I can add heat without necessarily increasing temperature: the energy could go into another form - chemical potential energy for example. Both Q and W are positive when energy flows into the system and negative when energy flows out.

# 13.6 The Formal Statement of the Conservation of Energy as the First Law of Thermodynamics

In section 13.4, we stated that energy must be conserved: . Moving things around we get . Recognizing the term on the left as , we can say . If we redefine and as just different directions of transferred energy , then we have where is positive if energy comes into the system and negative if energy is leaving the system. Now, we know that there are two different ways to transfer energy into or out of a system: heat and work. Thus, must be the sum of the energy transferred by heat and the energy transferred by work, . The statement of the law of conservation of energy can therefore be written as

.

Written in this form, the law of conservation of energy is called the *First Law of Thermodynamics*, i.e. the First Law of Thermodynamics and the Law of Conservation of Energy are the same thing.

This statement is so fundamental to the idea of physics that it is worth spending a minute to really unpack what it says. Looking again at the First Law of Thermodynamics (with the delta expanded) we see

where both heat and work are ways of *transferring* energy into or out of the system. As a first example, say we have some system and we do work on that system without transferring any energy as heat. In this case, as energy is coming in and . The result is that , which makes sense as we have added energy. Similarly, if we had a system that is losing heat to its environment while remaining stationary at constant volume then we know that because heat is flowing out and due to the fact that there is no “distance” for . Therefore, and as expected given that energy is flowing out of the system.

# 13.7 Why the First Law of Thermodynamics May Look Different in Your Other Courses

In some other courses or references, you may see the first law of thermodynamics written as , i.e. the sign of work may be different. This is still the same First Law of Thermodynamics/Law of Conservation of Energy that we are talking about here. The difference is one of perspective. In this class, we are considering energy flowing into the system as positive and energy flowing out of the system as negative. This convention matches our convention for heat as well as matching our definition of work from mechanics which considers only external forces. Physically, we are thinking about work done *on* the system by *external* forces.

To understand the formulation, you need a bit of history. The Laws of Thermodynamics were formulated during the Industrial Revolution as people were studying the properties of steam engines and the like. When studying the performance of a steam engine, the interesting quantity is not the work done *on* *the system by* *external* forces, but instead the work done *by the engine on its environment*. Stated another way, the developers of the Laws of Thermodynamics were not using our idea of object egoism! Instead of thinking about , they were thinking about . Now by Newton’s Third Law, these two forces are equal except for a negative sign. Thus, when you think about work done *by* the engine instead of the work done *on* the system, work flips sign and you end up with instead of .

**In this class, we will stick with , i**.e. we will use the same definition for work we have been using. The takeaway from this section is that you may see the First Law of Thermodynamics written with a different sign for work. Different fields use different conventions (it would be nice if we could agree, but oh well). Therefore, you should be aware that writing it as is just a different perspective born out of the historical development of science. This quirk with the sign of work is a great example of the impact that history and all of its associated socio economic factors can have on the history of science. One wonders what other ideas could be expressed more coherently? What scientific questions have not been explored because the people in power doing the research did not value them?

# 13.8 Enthalpy

*The following is a modification of content from 5.3 in OpenStax Chemistry*

Chemists and biologists often use enthalpy (*H*) to describe the thermodynamics of chemical and physical processes instead of the energy (*E*) and you may have seen this quantity before. Both enthalpy and energy have the same units and the words are similar so students often confuse these two ideas. However, while enthalpy and energy are related, enthalpy is not the same thing as energy. Energy is the ability to do work. Enthalpy is defined as the sum of a system’s microscopic or internal energy (*E*) and the mathematical product of its pressure (*P*) and volume (*V*):

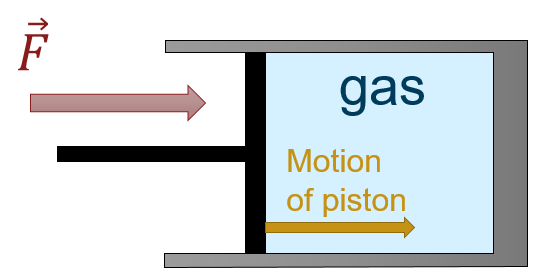
*H*=*E*+*PV*

Enthalpy values for specific substances cannot be measured directly; only enthalpy *changes* for chemical or physical processes can be determined. For processes that take place at constant pressure (a common condition for many chemical and biological processes due to the fact that they are open to the air which is at a constant pressure of 1 atm), the enthalpy change (Δ*H*) is:

Δ*H*=Δ*U*+*P*Δ*V*

Recall from the chapter on work that the mathematical product *P*Δ*V* represents work (*W*). If I compress the a gas in **FIGURE**, we see that the work is positive (the applied force is in the same direction as displacement of the piston) and energy is flowing into the system. However, the volume of the gas is shrinking. This example illustrates the general concept that the arithmetic signs of Δ*V* and *w* will always be opposite:

*P*Δ*V*=−*W*



**FIGURE:** *A gas being compressed by a piston. The force on the piston is in the same direction as the piston’s motion implying positive work. However, the volume of the gas is shrinking meaning*

Substituting this equation and the definition of internal energy into the enthalpy-change equation yields:

Δ*H*=Δ*E*+*P*Δ*V*

Δ*H*=(*QP*+*W)*−*W*

Δ*H*=*QP*

where *QP* is the heat of reaction under conditions of constant pressure. **Thus, if a chemical, biological, or physical process is carried out at constant pressure with the only work done caused by expansion or contraction, then the heat flow (*QP*) and enthalpy change (Δ*H*) for the process are equal.**

This condition, while it may seem restrictive, covers a significant fraction of the situations you will encounter. For example, the heat given off when you operate a Bunsen burner is equal to the enthalpy change of the methane combustion reaction that takes place, since it occurs at the essentially constant pressure of the atmosphere. On the other hand, the heat produced by a reaction measured in a bomb calorimeter ([[link]](https://cnx.org/contents/85abf193-2bd2-4908-8563-90b8a7ac8df6@9.480:0d364b67-be96-44fc-bee5-a368a42c2c82@11#CNX_Chem_05_02_BombCalor)) is not equal to Δ*H* because the closed, constant-volume metal container prevents expansion work from occurring. Chemists and biologists usually perform experiments under normal atmospheric conditions, which results in a constant external pressure making *Q* = Δ*H*.

*Note to David: leave that link to the bomb calorimeter*

Chapter 14 - Energy of Objects as a Whole (Macroscopic Scale)

For starters, it is simpler to think about energy of objects as a whole at the macroscopic scale separate from the collective energy of the constituent molecules. This thinking is in line with the physics problem solving approach of starting simple and adding complications later. The next chapter deals with energy at the microscopic realm. We will get into connecting these two realms in class. We shall see that there are only specific ways of transferring energy between the macroscopic world and the microscopic world so separating these two regimes makes sense. As you know from the previous chapter, heat is the transfer of energy by microscopic collisions. *Thus, heat is really only important at the microscopic scale and will not be considered in this chapter.*

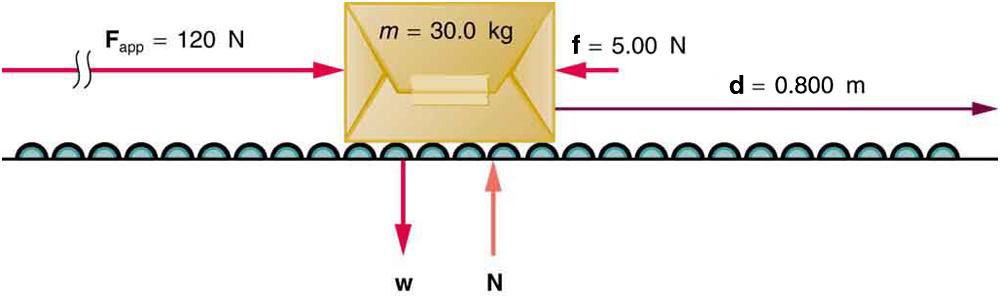
# 14.1 Kinetic Energy of an Object

*David: This section is modified from 7.2. I have made some edits throughout that I am highlighting in bold. This is due to the fact that we are starting from a place where the concept of conservation of energy is assumed (chapter 13).*

**Our goal in this section is to figure out an expression for the kinetic energy. While you should try to understand where the expression for kinetic energy comes from as it will help you understand how energy and work are related, the key is the expression we get to at the end.**

## 14.1.1 Figuring Out the Expression for Kinetic Energy

**To achieve this objective, let’s begin our study of energy with, as usual the simplest possible situation. Consider** a one-dimensional situation where a force is used to accelerate an object in a direction parallel to its initial velocity. Such a situation occurs for the package on the roller belt conveyor system shown in [Figure](https://cnx.org/contents/Ax2o07Ul@9.86:P_-6tVsN@5/Kinetic-Energy-and-the-Work-En#import-auto-id1803210).

* A package on a roller belt is pushed horizontally through a distance* ***d****.*

**In this case, there is no transfer of energy by molecular collisions, i.e. there is no heat and . Meaning that our statement of conservation of energy goes from**

**to**

**Similarly, there is no ability to do work due to position; the box cannot fall because of the rollers. Thus, there is no potential energy in this problem and all of our energy is kinetic energy: . Therefore, our statement of conservation of energy for this situation is just**

**.**

**The effect of the net force Fnet is to accelerate the package from *v*0 to *v*. The kinetic energy of the package increases, indicating that the net work done on the system is positive. (See** [**Example**](https://cnx.org/contents/Ax2o07Ul@9.86:P_-6tVsN@5/Kinetic-Energy-and-the-Work-En#fs-id1703845)**.) By using Newton’s second law, and doing some algebra, we can reach an expression for kinetic energy.**

The force of gravity and the normal force acting on the package are perpendicular to the displacement and do no work. Moreover, they are also equal in magnitude and opposite in direction so they cancel in calculating the net force. The net force arises solely from the horizontal applied force **F**app and the horizontal friction force **f**. Thus, as expected, the net force is parallel to the displacement, so that *θ*=0º and cos*θ*=1, and the net work is given by

*W*net=*F*net*d*.

**Substituting *F*net=*ma* from Newton’s second law gives**

***W*net=*mad.***

**To get a relationship between net work and the speed given to a system by the net force acting on it, we take *d*=*x*−*x*0**

**and use the equation studied in** [**Motion Equations for Constant Acceleration in One Dimension**](https://cnx.org/contents/031da8d3-b525-429c-80cf-6c8ed997733a@9.86:ea2bb23c-4fce-4e9d-a46b-3754125da988@9) **for the change in speed over a distance *d* if the acceleration has the constant value *a*; namely, *v*2=*v*02+2*ad* (note that *a* appears in the expression for the net work). Solving for acceleration gives *a*=*v*2−*v*022*d*. When *a* is substituted into the preceding expression for *W*net, we obtain**

***W*net=*m*(*v*2−*v*022*d*)*d*.**

**The *d* cancels, and we rearrange this to obtain**

***W*net=12*mv*2−12*mv* 20.**

## 14.1.2 Interpreting the Result: Kinetic Energy

*David: This section is taken from* [*http://umdberg.pbworks.com/w/page/68405433/Kinetic%20energy%20and%20the%20work-energy%20theorem%20(2013)*](http://umdberg.pbworks.com/w/page/68405433/Kinetic%20energy%20and%20the%20work-energy%20theorem%20(2013)) *with a few edits*

What has come out after all our manipulations is a that the work in this case is related to a change in a quantity associated with motion, ½*mv*2. This is kind of like momentum in that it counts both the mass and the velocity, but it differs in that momentum is proportional to the velocity vector -- so it is very directional. Reversing momentum is a big deal even if the speed doesn't change. For our new quantity, since it is proportional to *v*2 instead of to *v*, the direction of motion doesn't matter. You get the same *v*2 whether *v* is positive or negative. If our general result turns out to only depend on the magnitude of *v* and not the direction (it will), we will have solved our problem and learned what it is that changes an object's speed (not caring about direction).

When you compare the result of our manipulations to our analysis in terms of energy, you can see that ½*mv*2 must be the ***kinetic energy*.** It is a measure of "the energy associated with how much an object is moving".

# 14.2 Examples Applying Conservation of Energy with only Kinetic Energy

## 14.2.1 Calculating Kinetic Energy

*Use “Calculating the Kinetic Energy of a Package” example from 7.2*

## 14.2.2 Kinetic Energy Depends on the Square of the Velocity

*Make this look like the rest of the example problems.*

### Problem

A car travels at 5m/s when it accelerates to 10m/s. After the car has finished accelerating, by what factor did its kinetic energy increase?

### Solution

We are interested in the ratio of the final to the initial kinetic energies

Substituting our definition for kinetic energy we get

Since the mass of the car does not change, and the cancel leaving

Substituting our values in we get

### Analysis

The speed went up by two, but the kinetic energy went up by a factor of four, a result consistent with the fact that kinetic energy depends upon the square of the velocity. Speed matters a lot when thinking about energy!

## 14.2.3 Application of Conservation of Energy with Only Kinetic Energy

*Use a video of Heath’s that I am editing here*

<https://youtu.be/AEQUryH-twA?list=PLdIHHDayjYIK0G-CbsXZFt_JvX37tuAMD>

(should be under 5min)

# 14.3 Macroscopic Potential Energy

## 14.3.1 Gravitational Potential Energy

*For this use section 7.3 through the Kangaroo example. Swap PE → U and KE → K throughout. Also be sure to remove the parenthetical reference to the work energy theorem in the first paragraph of* ***Converting Between Potential Energy and Kinetic Energy***

## 14.3.2 The Zero of Gravitational Potential Energy

<https://youtu.be/WSsO1vs-dpY?list=PLdIHHDayjYIK0G-CbsXZFt_JvX37tuAMD>

## 14.3.3 Potential Energy of a Spring

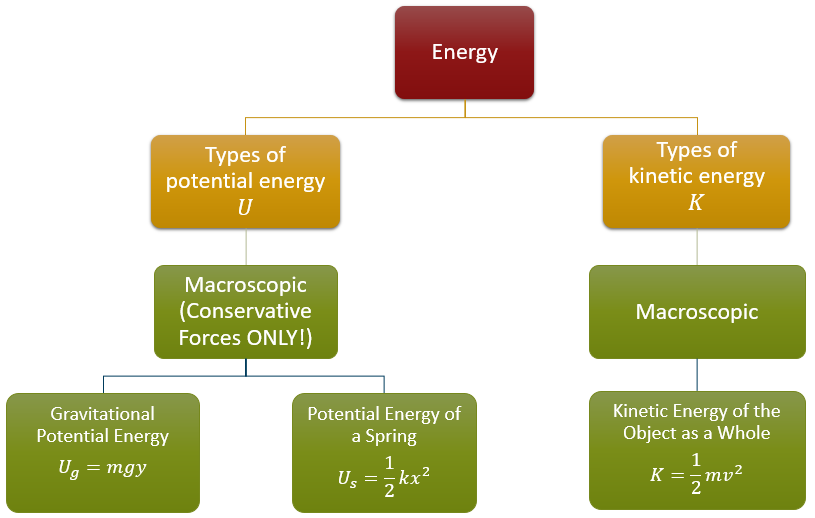
*For this use the* ***Potential Energy of a Spring*** *subsection of 7.4. Again make sure to swap PE → U*

# 14.4 Conservative vs. Non-Conservative Forces

<https://youtu.be/BmbLjnU77z8?list=PLdIHHDayjYIK0G-CbsXZFt_JvX37tuAMD>

# 14.5 Organizing the Different Types of Macroscopic Energy (Mechanical Energy)

Now we have all of the different types of macroscopic energy that we will talk about in this course: kinetic energy , gravitational potential energy , and the potential energy of a spring . These different types of energy can be organized as in the chart in the **FIGURE**. Collectively these types of energy are called *mechanical energy.*

****

***FIGURE:*** *The different types of macroscopic (also called mechanical) energy*

Using these forms, and noting that heat is a microscopic process, we can write the law of conservation of energy as

.

Where the work is the energy transferred into or out of the system by non-conservative forces. Knowing that the total energy is the sum of potential and kinetic energies, we can say that

which many people rearrange to look like

as this formulation separates the initial and final conditions. Finally, we could replace the potential energy with the types of potential energy are possible at the macroscopic scale: gravitational and springs. Such a replacement would leave us with

where is, for example, the initial gravitational potential energy.

This particular final form is very useful for analyzing situations in terms of energy. All we need to do to get started is:

1. Identify what types of energy we have in a given situation
2. Substitute the expressions for the gravitational potential energy , spring potential energy , and kinetic energy
3. Think about any energy entering or leaving the system as non-conservative work

We will look at applying these ideas in class. The next section has some examples of this process for the special case where the energy entering or leaving the system as non-conservative work is zero; situations where the mechanical energy is conserved. These are the types of problems that we will expect you to be able to do on your homework and quiz. We will address the situation where energy is entering or leaving the system in class.

# 14.6 Examples where Mechanical Energy is Conserved

<https://youtu.be/8SrC6hiOodYv>

## 14.6.1 Kingda Ka

<https://youtu.be/Zrp-W5GM_58?list=PLdIHHDayjYIK0G-CbsXZFt_JvX37tuAMD>

## 14.6.2 Wrecking Ball

<https://youtu.be/S40uRk8ENU8?list=PLdIHHDayjYIK0G-CbsXZFt_JvX37tuAMD>

## 14.6.3 Example with a spring

<http://umdberg.pbworks.com/w/page/89285639/Energy%20conservation%3A%20Example%201>

# 14.7 Graphical Tools for Analyzing Situations Using Energy

As we have seen throughout this course, being able to represent concepts in multiple formats can be useful. Throughout the course, we have represented ideas in words, equations and graphs. Graphical representations can also be a useful way of analyzing situations involving energy, albeit we will need a new type of graph: a pie chart.

Pie charts are particularly useful in the, common, case where there is little to no energy entering or leaving the system of interest. In these cases, the total amount of energy is constant and the only changes are from one form of energy to another. The quantity of interest is then: what fraction of the total energy is in each form? Representing such fractions of a whole is the strength of the pie chart.

## 14.7.1 A First Example

### Problem

The use of pie charts is probably best illuminated by an example. As per the usual approach in physics, we will start with the simplest possible case to get the fundamental principles and then add complications. Consider the case of a ball being dropped straight down as in **FIGURE**. We will only look at the situation before the ball hits the floor and ignore any impacts of air resistance. Let’s think about the energy for the four instants shown:

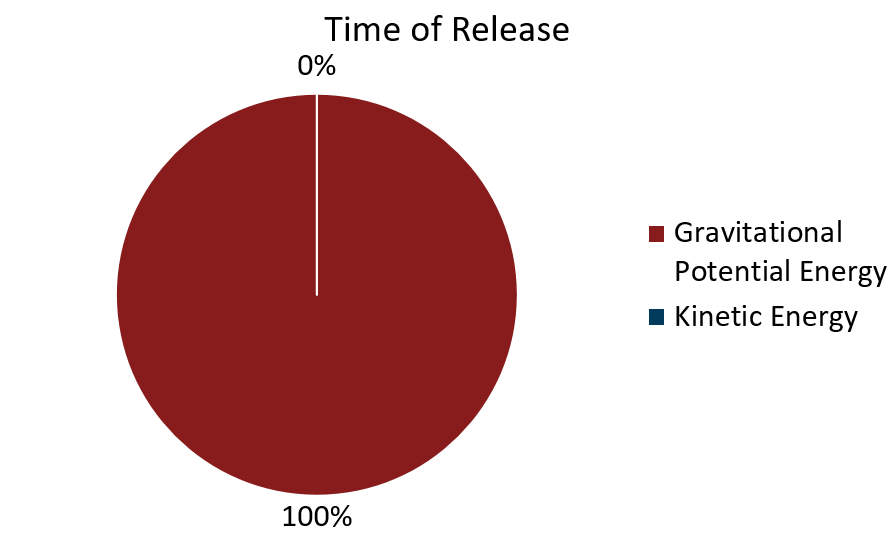
1. The instant the ball is released
2. A short time after the ball
3. Halfway to the ground
4. Just before the ball hits the ground



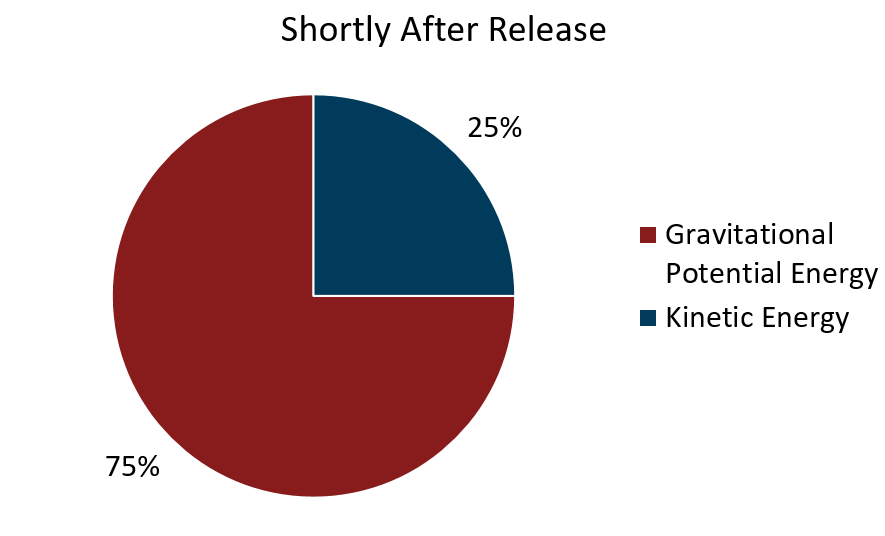
***FIGURE:*** *A ball falls from a height. Four instants in the fall are shown: the ball as it is just released, the ball after falling a short distance, the ball after falling ½ the way, and the ball just before it hits the ground. The arrows to the side get longer to represent that the ball is moving faster as it falls.*

### Solution

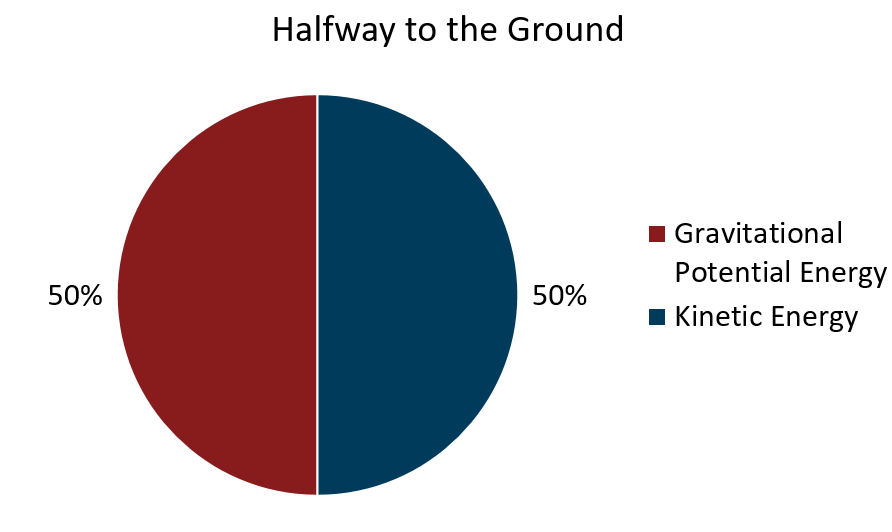
1. The instant the ball is released the ball is not moving at all and therefore has no kinetic energy. All of the ball’s energy is gravitational potential energy. We would represent this with a simple pie chart showing 100% of the energy as gravitational potential energy



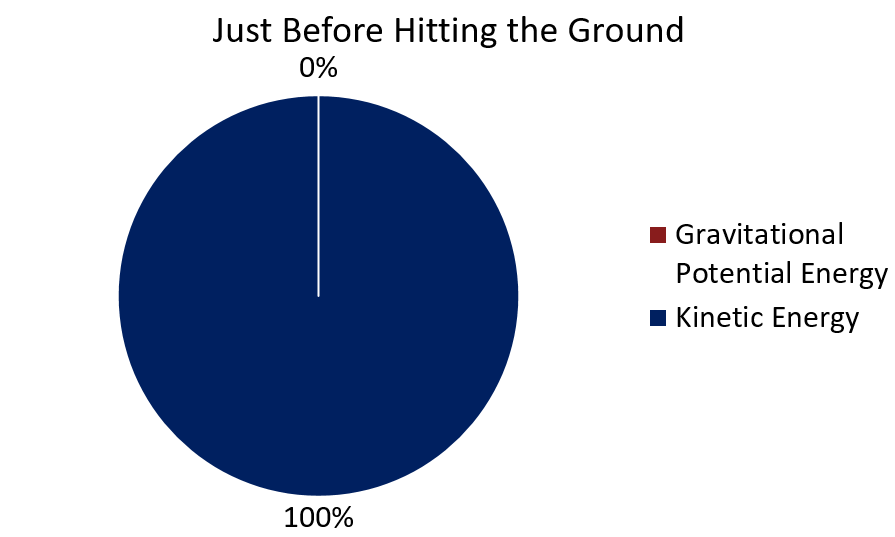
1. A short time after the ball is released, the ball has gained a little bit of kinetic energy, but most of the energy is still gravitational potential energy. Now, not enough information was provided to be quantitative, but we can say that most of the pie should still be gravitational potential energy and our graph might look like the one below.



1. When the ball is halfway to the ground we can be a bit more specific. Initially the ball was at some height , and therefore had an initial gravitational potential energy . Now that the ball is halfway to the ground, the height must be and the gravitational potential energy is or ½ of its initial value. Since the total energy must be constant, the rest of the energy MUST be kinetic and we end up with a pie chart that is 50% potential energy and 50% kinetic energy as shown.



1. At the instant before the ball hits the ground, the height of the ball is essentially zero and therefore the ball has no more potential energy and all of its energy must be kinetic. Therefore, at this instant, we get a pie-chart that is 100% kinetic energy as shown.

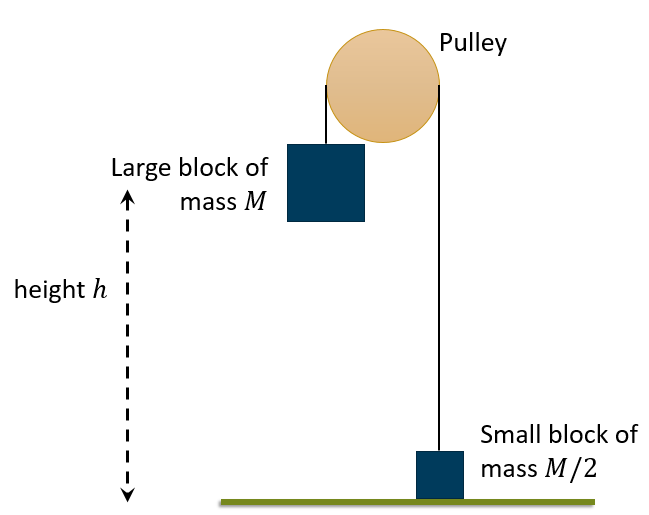


## 14.7.2 Another Example with Two Objects

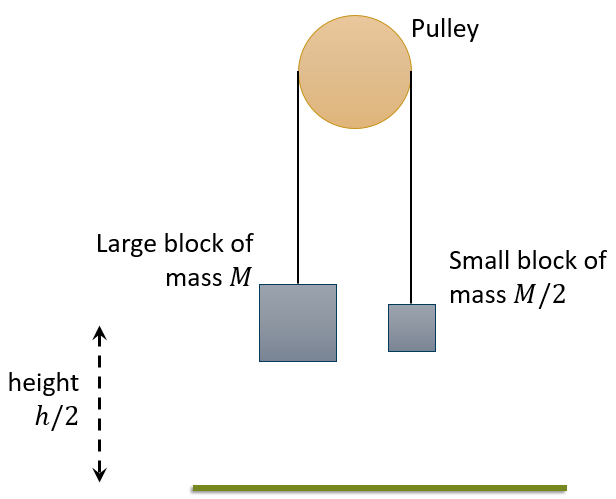
### Problem

In this case, we have two objects of unequal mass tied together over a frictionless pulley as shown. The heavier object is suspended above the ground on which the lighter object, which has half the mass rests. Draw the energy pie charts for:

1. The initial condition
2. The situation where the large mass has fallen halfway to the floor and the small block has, due to the fact that they are tied together, risen the same distance.

**

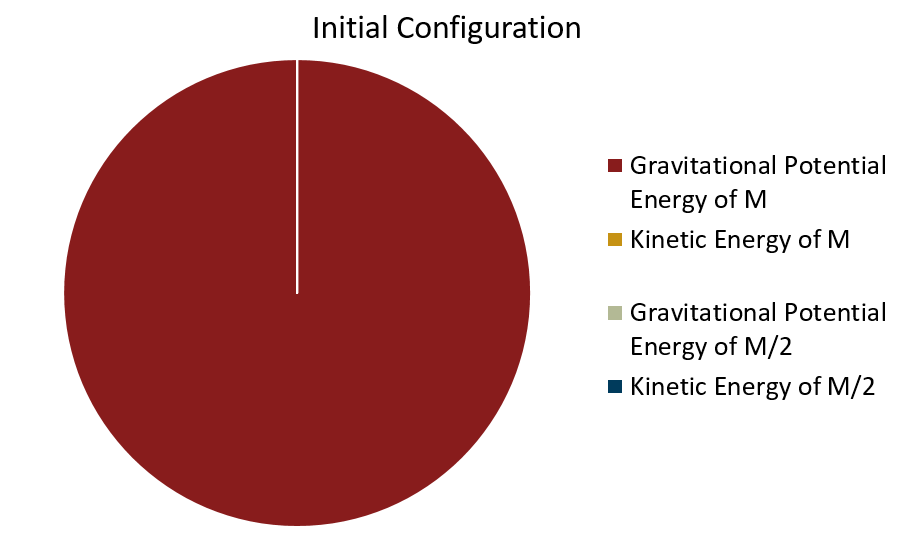
***FIGURE:*** *The setup for this example as described in the text where the large block of mass is suspended a height above the ground on which a small block of mass rests.*

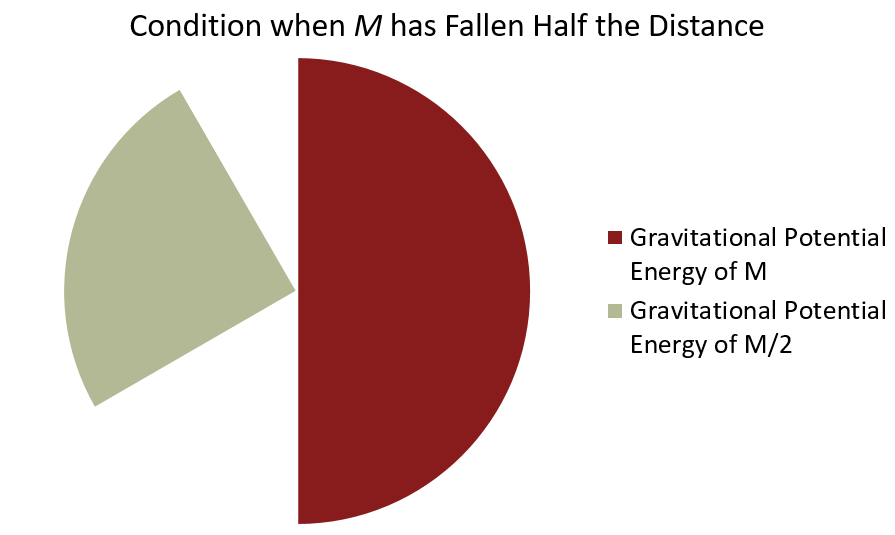
**

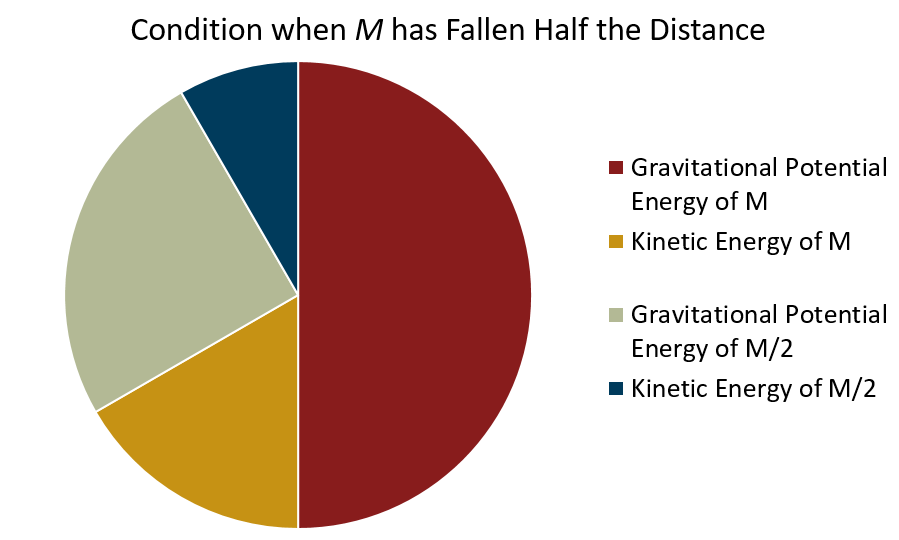
***FIGURE:*** *The final condition of the situation of interest, the large block has fallen half of its initial height while the small block has, because they are tied together, risen the same distance.*

### Solution

1. For the initial conditions, all of the energy is in the gravitational potential energy of the large block. The total energy in the system is



1. This situation is a bit more interesting and it would be best to look at the different elements of the system one at a time. Remember the total energy of the system cannot change and is   
     
   Potential Energy of : The block is now at which means that its potential energy is now which is half the total energy of the system so the potential energy of should be 50% of our pie.   
     
     
   Potential Energy of : This block has risen (the distances must be the same as the two blocks are tied together). Thus, the potential energy of this block is now or ¼ of the total. Thus the potential energy of should be 25% of our pie.   
     
     
   This leaves the two kinetic energies, which by conservation of energy MUST add up to 25% of the total. Moreover, the two blocks MUST be going at the same speed (again, because they are tied together). However, the two blocks do NOT share the kinetic energy equally. The block is twice the mass of the block ! If the two kinetic energies add up to 25%, then we have   
   .  
   Knowing that and substituting we can see   
     
     
     
   or should be 16.67% of our pie! Finally, must be half of that or 8.33%. The result is the final pie shown below



## 14.7.3 How Energy Pie Charts can be Used to Analyze Situations (Take-away)

Energy pie-graphs are a useful tool to analyze situations in terms of energy. The pie graphs can be used to:

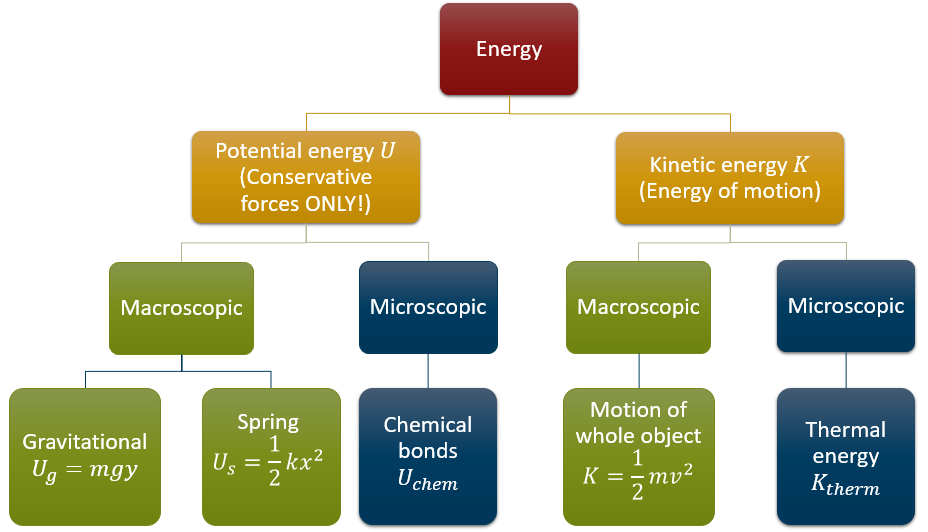
* Make sure that you have taken all possible types of energy into account
* Get a feel for how much of the total energy is being carried in each form
* Visualizing the transformation of energy from one form into another

Conservation of energy only holds if you take all the different forms of energy into account. Thus, making sure that you have everything is critical, as you have seen in the algebraic examples solved in section 14.6. Pie graphs are a good way to separate the analysis and just think about the types of energy you have before getting into the math. At the same time, getting a feel for what is going on which can allow you to check your result.

Chapter 15 - Energy of Constituent Atoms (Microscopic Scale)

In this chapter, we will be exploring the idea of energy at the microscopic scale. Instead of talking about the energy of the object as a whole we will be talking about the kinetic and potential energies contained within the molecules themselves. This energy is generally MUCH larger than the mechanical energy at the macroscopic scale and is of fundamental importance to our modern world and the subjects of biology and chemistry which are greatly concerned with the conversion of microscopic chemical and thermal energy into useful work. This chapter will ONLY deal with the microscopic world; just as the last chapter dealt solely with the macroscopic realm. We will look at how to connect these two different distance scales in class.

In this chapter, we are still dealing with the First Law of Thermodynamics where the total energy is still the sum of the potential energies and kinetic energies : . The only difference is that now the work as well as the types of potential and kinetic energies will be microscopic. Since we are looking at the microscopic scale, heat will play more of a role (recall that heat is the transfer of energy by *microscopic* collisions!) than it did at the macroscopic scale of the last chapter. The form of potential energy that we will be mostly concerned with at this scale is the potential in molecular bonds: so-called chemical potential energy . In terms of kinetic energy, we shall see in this chapter that kinetic energy at the microscopic scale is related to the temperature of the object: is related to . The relationships between these different forms of energy are shown in **FIGURE**. The total energy at the microscopic scale is sometimes called the *internal energy* of the system.



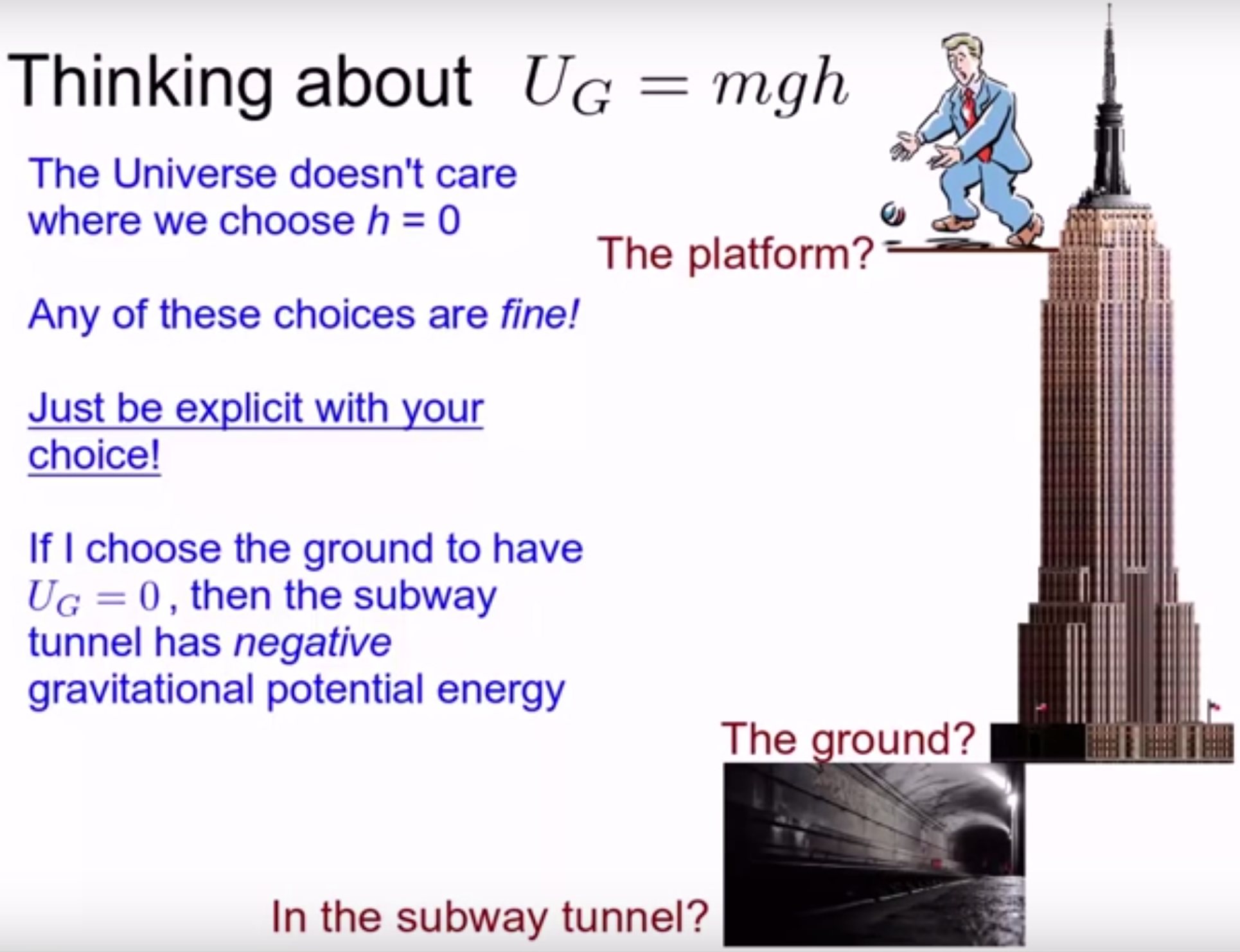
***FIGURE:*** *The different types of energy classified as microscopic and macroscopic*

# 15.1 Potential Energy of Molecules

As stated in the introduction, the primary source of microscopic potential energy with which we shall concern ourselves is the potential energy stored in chemical bonds or chemical potential energy . This potential energy is a result of the force of electrical attraction between different atoms (recall electricity and magnetism was one of our fundamental forces). As you shall see in the next course, the electrical force is a conservative force and thus we can associate a potential energy with it.

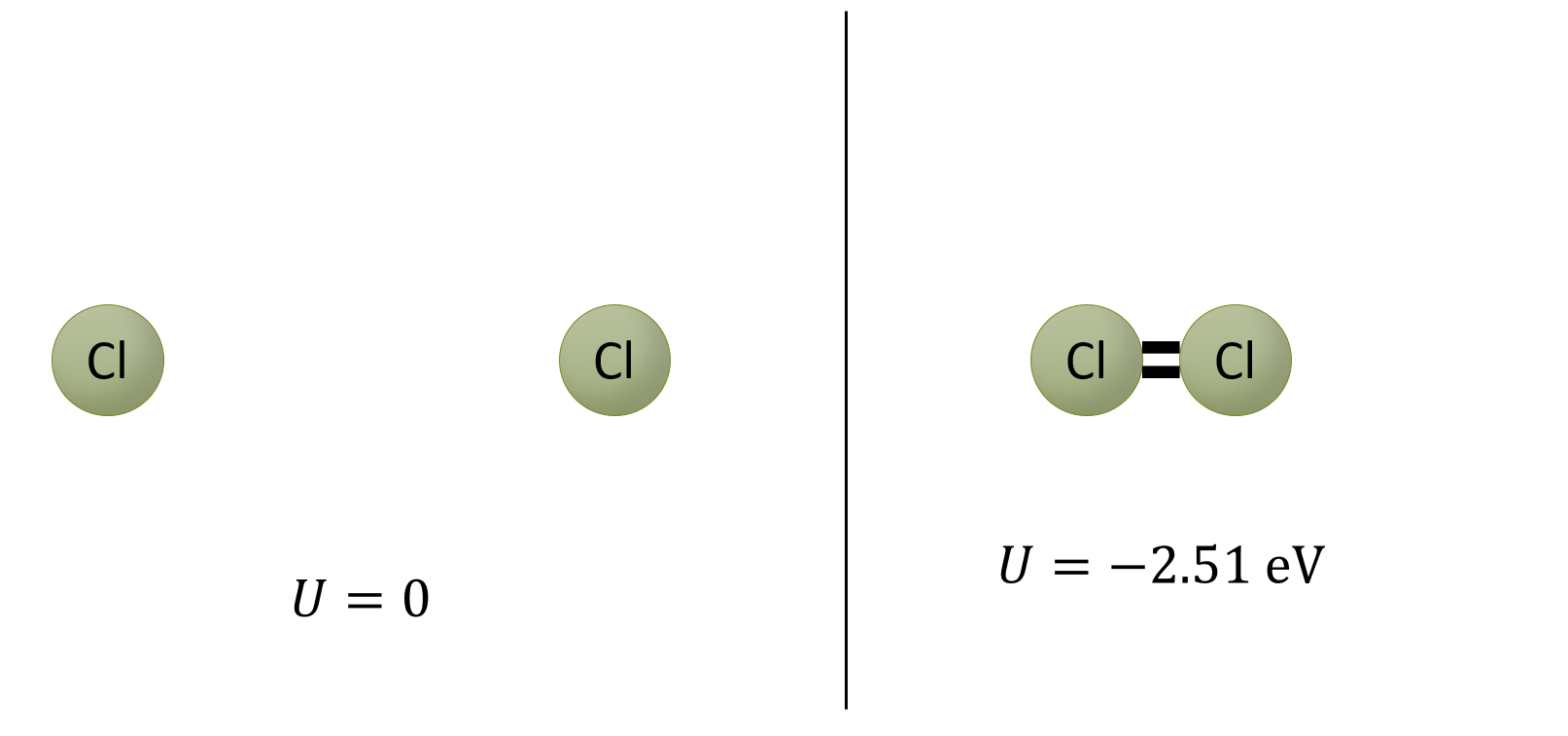
The strength of chemical bonds is typically quoted in one of two ways: either the energy in the bond, called the *bond dissociation energy*, is quoted directly (typically in eV) or the enthalpy per mole will be quoted. For example, the Cl-Cl bond has a bond dissociation energy of 2.51 eV/bond or a bond dissociation enthalpy per mole of . How do we interpret these numbers in terms of potential energy? We use the same freedom to choose the zero of potential energy that we discussed in section 14.3 when we discussed potential energy at the macroscopic scale.

Thinking about gravity, we tend to put the zero of potential energy at ground level; objects above the ground then have positive potential energy while objects underground have negative potential energy. This use of negative potential energy makes sense, an object at ground level will fall to below ground level if allowed to do so and lose energy in the process.

****

***FIGURE:*** *A clip from the video on zero of potential energy showing that negative potential energies are possible!*

For atoms and molecules, we have a similar freedom to choose where to put zero potential energy. The standard convention is to say that free atoms that are far apart have *zero* potential energy. Atoms in most bonds have lower potential energy than free atoms (that is why the bonds form!). Therefore the potential energy of the atoms is less than zero: the potential energy of atoms in bonds is *negative*. This may seem like a weird choice for the zero of potential energy, but it is the convention and it makes sense when you think about it!



***FIGURE:*** *Two Cl atoms separated by a great distance have zero potential energy while two bonded Cl atoms have a potential energy of -2.51 eV. Remember, potential energy is the due to the relative position of two objects, so it does not make sense to ask which atom in the bonded pair has the potential energy. The potential energy is due to the two of them!*

Let’s return to the quoted Cl-Cl bond with dissociation energy of 2.51 eV/bond. What does this value mean? It means that two Cl atoms bonded together have a potential energy of 2.51 eV *less* than if they were free. Said another way, the potential energy of Cl atoms in Cl2 is -2.51 eV, while the potential energy of free Cl atoms is 0 eV. This is consistent with what you probably already know about Chlorine: Cl2 is the lower energy state than free Cl atoms. I would get 2.51 eV of energy for every Cl-Cl bond that is formed, as the atoms move from zero potential energy to -2.51 eV. Similarly, I would need to add 2.51 eV of energy to break a Cl-Cl bond and move the two atoms *up* to zero potential energy.

# 15.2 Application of Bond Energies

Let’s see how we can use these two different ways of quoting chemical energy and convert the result into something we can use.

## 15.1.1 Example: Bond energy expressed in eV

Back in chapter 13, we saw that the bond dissociation energy of the H-H bond was 4.52 eV/bond. How much energy is released by converting 6.000g of monotomic H into 6.000g of H2?

### Solution

Also at <https://youtu.be/jBQGwjokX14>

We have 6.000g of monotomic hydrogen. The molar mass of hydrogen is 1.008g/mol. Therefore we have

Each molecule of H2 requires two H atoms so we can make

We know, from the table that each bond releases 4.52 eV, so

which we can convert to Joules

## 15.1.2 Example: Bond energies expressed as enthalpies

The enthalpy of dissociation of CO2 → CO + O is . How much *energy* is required to break up 20g of CO2 into CO and O assuming the reaction occurs at constant pressure?

### Solution

Also at <https://youtu.be/SE_SN3plzbY>

We need to convert from enthalpy *H* into energy. The definition of enthalpy is (from section 13.8)

At constant pressure, the change in enthalpy is

Using the first Law of Thermodynamics

Recalling from section 13.8 the argument that, under constant pressure, the work and are of equal magnitude, but opposite sign, we have

(This is the exact same reasoning we did in 13.8 to show that, under constant pressure , we are just repeating the steps to reinforce the logic).

Now, we know that the *enthalpy* change is the same as the *heat* . Since the heat is positive we must add heat to make the process go; one mol of CO2 going to CO and O will need 532 kJ of energy.

**Instructor’s note:** There are three different quantities in that last sentence: enthalpy, energy, and heat. Furthermore, all three have units of Joules! Make sure you understand the distinctions between these three concepts!

Now, all we need to do is determine the amount required to break up the 20g of CO2 in the problem. From the periodic table we see that the molar mass of CO2 is:

Now to get the total energy that we must add as heat:

We must add 241 kJ to break up 20g CO2 into CO and O.

## 15.1.3 Example: Using chemical potential energy in conservation of energy

The bond dissociation energies for various stages of the 2H2 + O2 → 2H2O reaction are shown in the table below. How much energy is released by converting 2 molecules of H2 and 1 molecule of O2 into 1 molecule of H2O at constant pressure?

|  |  |
| --- | --- |
| Bond | Bond energy (eV/bond) |
| H-H | 4.52 |
| O=O | 5.15 |
| O-H | 4.77 |
| H-OH | 2.78 |

### Solution

We know from experience that we can burn H2 and O2 and get water and energy out. Thus, we know that in the law of conservation of energy:

The heat will be negative (recall from 13.5 that heat leaving the system is negative!).

Initially, we have two H-H bonds and one O=O bond:

Recalling that bond energies should be negative

At the end we have two O-H bonds (one for each molecule) and two H-OH bonds

We see that the final energy is less than the initial, which matches our expectations! The total heat released can now be calculated

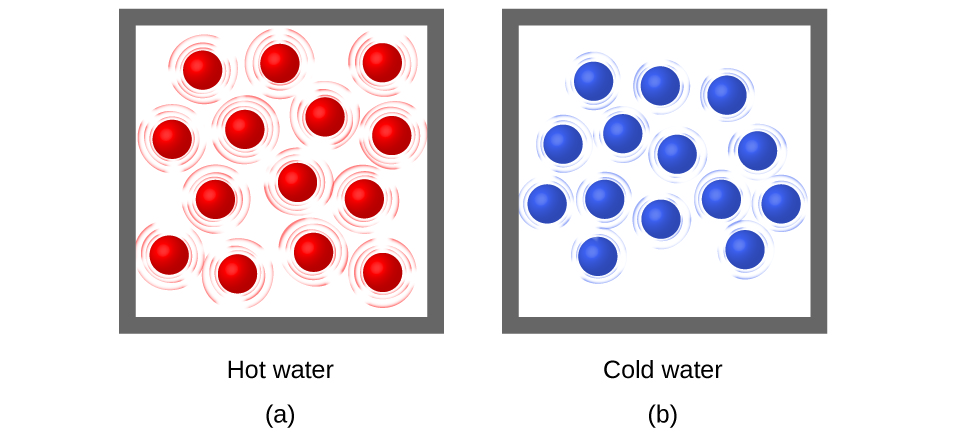
This reaction will release 7.43eV of energy. Mathematically, we know that the energy is released because the heat is negative implying energy leaving the system.

# 15.2 Kinetic energy at the Microscopic Scale

Now, we have talked about potential energy at the microscopic scale. What about kinetic energy? Back in chapter 12, you were refreshed on the idea of temperature. What is temperature *really*? It turns out that the ideas of temperature and the kinetic energy at the microscopic scale are deeply related. Along the way, we will introduce the idea of *degrees of freedom* or places to put energy. This concept will also be important in our next unit on entropy.

## 15.2.1 Microscopic Temperature

So what is temperature? Temperature is, in essence, a macroscopic measurement of the average kinetic energy of molecules - a microscopic quantity. Molecules are always moving and vibrating around. The more kinetic energy the molecules have, the higher their temperature as shown in **FIGURE.**

****

***FIGURE:*** *(a) The molecules in a sample of hot water move more rapidly than (b) those in a sample of cold water*

However, there is a bit of a complication. When I add energy to a molecule I can put it in a bunch of different places or *degrees of freedom*. Temperature is then formally related to the average energy per degree of freedom. In equation form

or more commonly written

.

In this expression is the average energy, is the number of degrees of freedom (places to put energy), and is the temperature *in Kelvin*. The factor of comes out of a formal calculus-based derivation, but as far as we are concerned, it is just there for convenience purposes. The symbol is the Boltzmann constant introduced in 12.4 and appears anytime you try to connect the microscopic world to the macroscopic world. In this case, we are connecting microscopic kinetic energy, measured in Joules, to the macroscopic quantity known as temperature; thus the units J/K. We will see this constant again playing a similar role connecting the microscopic and macroscopic worlds in our discussion of entropy in the next unit.

## 15.2.2 Degrees of Freedom

What do we mean by a “place to put energy?” Consider a simple ideal gas molecule. Now add some energy to it. Where can that energy go? Well we know from section 3.1 that the *x*, *y*, and *z* directions are independent of each other. Thus, the energy I add could go into either kinetic energy associated with the *x*-direction, kinetic energy associated with the *y*-direction, or kinetic energy associated with the *z*-direction as shown in the **FIGURE**. Thus, an ideal gas molecule has *three degrees of freedom*. Remember, kinetic energy itself does not have direction. We are just talking about different places *on each atom* that we could put that energy - a different concept. We will talk about counting degrees of freedom for other gases in class.

*David: Append* <https://youtu.be/QXjZpUDIC6A>

## Example using temperature, kinetic energy, and degrees of freedom

(a) What is the average kinetic energy of an ideal gas molecule at 20.0ºC (room temperature)? (b) Find the average speed of a helium atom (He) at this temperature.

### Solution for (a)

We know that the average kinetic energy of a molecule is related to the temperature by the expression

For an ideal gas molecule, the number of degrees of freedom is 3 so we have

Before we can use this expression, we need to convert the temperature to Kelvin

Now substituting it all together we have

### Solution for (b)\*

Now we know the average kinetic energy, we can think about the average speed. We begin with the definition of kinetic energy

.

Both ½ and are constants and can come out of the average

>

The mass of helium is, from our periodic table, 4.002 amu which is . Putting this value and our result from (a) in our expression we get

Pretty quick!

\*In this calculation, we actually calculated not the average speed but the root-mean-square speed. This calculation is pretty much exactly what it says: we calculated the average of all the velocities squared (square each, add them up, divide by the number) and then took the square root. The average velocity of the He atoms is zero (velocity is a vector so moving left cancels moving right)! We will ignore this distinction in this class but it may be important in a future class!

# 15.3 The Total Kinetic Energy in and Object Associated with its Temperature

We now know that temperature is the average kinetic energy in an atom per degree of freedom (with a factor of 2 and a thrown in). How much energy in total is in an object just due to the kinetic energy of all of its molecules wiggling around due to their temperature? If the average energy per atom is

,

then to get the total kinetic energy due to temperature , we take the average and multiply by the number of atoms

.

Making a substitution we have

.

Dealing with numbers of molecules can be unwieldy, so let’s convert to moles . You will recall from section 12.4 that

where is the gas constant 8.314 J/(mol∙K), meaning we can rewrite our expression as

## Example: Compare total thermal energy to gravitational potential energy

Consider a block of lead with about the same mass as a person, 70kg, sitting on the second floor of a building, 3m above the ground, at room temperature of 20oC. What is the ratio of gravitational potential energy to the total thermal energy due to the block’s temperature?

### Solution

The gravitational potential energy of the block is given by

The thermal energy is given by

Since lead is a 3-D solid, we know that the number of degrees of freedom is 6. To get the number of moles, we look up the molar mass of lead and find it to be 207 g/mol which means that 70kg of lead is

Putting it all together (and converting 20oC to Kelvin) we get a total thermal energy of

Giving us a ratio of

i.e. the thermal energy is over 1000x bigger than the gravitational potential energy!

### Analysis

We see that the internal thermal energy in this case is MUCH bigger than the gravitational potential energy, a common result. This disparity is why, when considering the energies at the microscopic scale, the macroscopic energies can usually be ignored. However, there are situations where energy moves from one scale to another such as a car engine which converts microscopic potential energy in the gasoline into microscopic kinetic energy of hot gas in the engine, into motion of the car. Clearly in these situations, you need to think about both the microscopic and macroscopic scales.

## 

Say we have a container with 0.25mol of argon gas which we can treat as being an ideal gas. How much energy do we need to add as heat to raise the temperature 10oC assuming no work is done?

### Solution

We begin with the First Law of Thermodynamics (Conservation of Energy)

which, given that no work is done goes to

.

All of the energy in this case is thermal kinetic energy. Replacing with , therefore, we have

.

Now, we put in the definition of total thermal kinetic energy to get

which after factorization becomes (note this equation for the next section!)

.

Argon is an ideal gas, so . We know the number of moles and is the gas constant . Moreover, we know the temperature difference . We, as always, need to convert to Kelvins. However, the size of is the same as , only the zero points are different. Therefore . Substituting these values in, we get

The result is positive, so we need to add 31J of heat energy to the system to get the temperature to increase 10oC (which makes sense intuitively).

# 15.4 Connection to Chemistry

**Instructor's Note: This section is a bit of a derivation, which we try to avoid. However, in this case it is worth it. Spend some time trying to follow along; get out a piece of paper and work it for yourself. The result will be a “woah” moment that explains numbers you had to look up in a table in chemistry!**

If you have taken chemistry before, then the context of the previous problem may seem familiar to you. Problems about how much energy you need to raise or lower the temperature of a substance are often described in chemistry books as being calorimetry problems *(David provide a link to OpenStax Chemistry section on calorimetry)* to be solved using the expression

.

In this expression, is the mass of substance, is the change in temperature, and is a quantity called *specific heat* with units . If you have done these types of problems, then you have probably either solve for the specific heat or looked it up in a table like the one below for the substance in which you were interested.

|  |  |  |
| --- | --- | --- |
| Substance | State | Specific Heat at constant volume (J/gK) |
| Helium | He(g) | 3.12 |
| Water | H2O(l) | 4.19 |
| Ethanol | C2H6O(l) | 2.3 |
| Nitrogen | N2(g) | 0.743 |
| Aluminum | Al(s) | 0.87 |
| Argon | Ar(g) | 0.315 |
| Iron | Fe(s) | 0.46 |
| Copper | Cu(s) | 0.39 |

***Table:*** *A table of specific heats measured while holding the volume constant taken from* [*http://www.engineeringtoolbox.com/specific-heat-capacity-gases-d\_159.html*](http://www.engineeringtoolbox.com/specific-heat-capacity-gases-d_159.html)

It turns out that, with what you know now, you can actually *predict* these specific heats to a high degree of accuracy based upon the properties of the molecule!

Let’s compare the chemistry calorimetry formula

with the equation you were asked to note in the last example

which after some rearranging becomes

.

In both expressions we have a quantity telling us how much material we have (mass or number of moles), multiplied by some number, and then multiplied again by the change in temperature to get the amount of heat.

To further see the comparison, let’s multiply our result by a clever form of the number 1, where is the molar mass of the substance in g/mol (I know this seems dumb, but bear with me for just a few lines):

This is one of a physicist's favorite tricks and is totally legit: I did the same thing to both sides. One the right-hand side of our equation, where the heat is, nothing happens, I multiplied by 1 after all. On the left-hand side, however, I can rearrange things in an interesting way:

.

All I did is take the top molar mass and put it next to the number of moles , and took the bottom molar mass and put it with the 2. Now, is the mass of the substance by definition; has units g/mol and is the number of moles leaving us with grams. Making this substitution we get

which when we compare with the calorimetry formula from chemistry

we see that the only way this can work is if

## Example: Let’s test this result for two materials

We shall test this result for two different materials (a) argon and (b) copper and compare with the table above

#### Solution for argon

Argon is an ideal gas so . The molar mass for argon is 39.948 g/mol. Substituting into our expression we get

which matches the table value exactly!

#### Solution for copper

Copper is a solid with and a molar mass of 63.546 g/mol. Thus, we expect the specific heat of copper to be

Again, a perfect match!

## Takeaway

This is our first example where microscopic properties of atoms can be used to get macroscopically measured phenomena: we went from the microscopic structure of atoms to predicting the macroscopic specific heat. We will do more of connecting these two worlds in class throughout this unit. The next unit on entropy continues this pattern.

1. “Bond-Dissociation Energy - Wikipedia.” Accessed August 1, 2017. https://en.wikipedia.org/wiki/Bond-dissociation\_energy. [↑](#footnote-ref-0)
2. *File:Food Label from The Co-Opertative Food Sausages.Jpeg - Wikimedia Commons*. Photograph, June 8, 2015. https://commons.wikimedia.org/wiki/File:Food\_label\_from\_The\_Co-opertative\_Food\_Sausages.jpeg. [↑](#footnote-ref-1)
3. SteveBaker. *English: A 2007 MINI Cooper’S Car Shown Immediately before - and Soon after - a Severe Car Crash. At: Lat/Long 31.03669,-97.470881*, February 9, 2009. Own work. https://commons.wikimedia.org/wiki/File:BeforeAndAfterMINICooperS.png. [↑](#footnote-ref-2)