

UMass Physics 131 Unit 4

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UNIT 4 OVERVIEW

This overview is based on

umdborg / *Reductionism and emergence* (2015). Available at: [http://umdborg.pbworks.com/w/page/68371403/Reductionism%20and%20emergence%20\(2015\)](http://umdborg.pbworks.com/w/page/68371403/Reductionism%20and%20emergence%20(2015)) ([http://umdborg.pbworks.com/w/page/68371403/Reductionism%20and%20emergence%20\(2015\)](http://umdborg.pbworks.com/w/page/68371403/Reductionism%20and%20emergence%20(2015))) . (Accessed: 23rd August 2017)

Science considers a wide variety of scales. In physics and biology we consider

- the nanoscale of atoms and molecules -- distances of nanometers (atomic size) and timescales of microseconds;
- the microscopic scale used in cellular biology -- distances of microns to millimeters and times of fractions of a second;
- the macroscopic human scale we are used to in everyday life -- distances of meters to miles and times of seconds to years;
- the global/deep time scale of ecology and evolution -- distances spanning thousands of kilometers and times of millions of years.

Physics considers scales beyond this including subatomic and even subnuclear structures and cosmological scales to the size of the entire visible universe and timescales of the lifetime of the universe.

We tend to analyze each system on its own terms using the concepts appropriate to the scale. But some of the most interesting scientific insights come from crossing scales. When we explain the properties of systems at one scale in terms of its component parts at a finer scale it's called **reductionism**. So when we analyze the conductivity of copper in terms of the bonding properties of copper atoms and how they share electrons from one atom in a crystal to the next, it's reductionism. When a patient experiences palpitations, sweating, dizziness, and headache, a caregiver might interpret this as occurring as a result of an imbalance in a chemical in the bloodstream -- insulin. Both of these explanations of a macro event in terms of microscopic properties are reductionism. We are looking at a system at a large scale and its characteristics are explained by something happening at a smaller scale.

We can consider the phenomenon of scale crossing in the other direction. If we are looking at a system at one scale, we might find that effects that seem very small add coherently to produce a dramatic and important effect when we step back and look at things at a larger scale. When the properties of a system at the scale we are considering have effects on a system of a larger scale, it's called **emergence**, especially when the phenomenon might be almost un-noticable at the scale we are considering.

A physical example is polarization. If we put an atom in an electric field of typical macroscopic values (a few volts per meter), the electrons and the nucleus are pulled apart -- but only by a very tiny amount, perhaps 1 part in 100,000 of the atom's diameter. We might assume this is so tiny an effect that it can be ignored; but if every atom in a macroscopic object undergoes the same slight separation, the total effect might be that we can pick up the object, lifting it against gravity with the sum of the tiny electric forces we are exerting on each atom. The fact that we have so many atoms multiplies what looked to be a tiny effect. A similar phenomenon occurs in biology when there is a small survivability advantage to a mutation. One might not see any effect for many generations, but given thousands of generations, the gene pool of a population can be completely transformed by natural selection.

Of course, these are the same phenomenon, just looked at from different angles. But whatever scale we are considering, keeping these two perspectives in mind will help us to look for structures both at smaller scales that might provide reductionist explanations and at larger scales that might have emergent properties.

1 INTRODUCTION TO THE MICROSCOPIC WORLD

1.1 Introduction

UMASS AMHERST Instructor's Notes

We expect that all of the material in this chapter is review for you from other courses, and is included here for reference. The topics with which you are least likely to be familiar are:

- The expression of the ideal gas law in terms of molecules instead of moles and k_B vs R in section 12.5
- The limitations of the ideal gas law in section 12.5
- The Einsteinian Solid in section 12.6

Your quiz will cover:

- Converting between atoms, grams, and moles for any compound given a periodic table.
- The limitations of the ideal gas law

This introduction is based from

umdborg / The micro to macro connection (2013). Available at: [http://umdborg.pbworks.com/w/page/68405576/The%20micro%20to%20macro%20connection%20\(2013\)](http://umdborg.pbworks.com/w/page/68405576/The%20micro%20to%20macro%20connection%20(2013)) ([http://umdborg.pbworks.com/w/page/68405576/The%20micro%20to%20macro%20connection%20\(2013\)](http://umdborg.pbworks.com/w/page/68405576/The%20micro%20to%20macro%20connection%20(2013))) . (Accessed: 2nd August 2017)

Much of what we have done so far -- the Newtonian framework, describing the properties of solids and liquids, and the concepts of heat and temperature -- are *macroscopic* concepts: they describe things we see, feel, and experience. They express the regularities and consistencies of the behavior of physical systems. Much of this was well known by the middle of the nineteenth century. But one of the most extraordinary and important pieces of knowledge that humanity has garnered since then is the idea of the *microscopic*. By this, physicists don't mean "what you can see in a microscope", but rather the fact that everything we regularly experience is made up of a small number of different kinds of atoms (91 in the natural world, a few more that have been created by humans). The essential point about this is that we believe that all properties of the macroscopic world are ultimately due to the properties and interactions of those 91 distinct elements. Although some phenomena require a description at a higher level (see the discussion of emergent phenomena), at some level (even if it's not convenient or useful for us to explicate), everything we see is a result of atomic properties.

A major component of modern biology is working at the microscopic -- atomic and molecular -- level and learning what are the critical elements that underlie basic biological mechanisms. Much of the research and development that can be expected to transform both biology and medicine over the next few decades will depend on making sense of the micro to macro connection. In this class, we will develop a few of the basic tools needed for making this connection. One set of tools involves *statistical physics*. Since there is a huge amount of energy distributed in all objects at common temperatures, and since these energies tend to be *randomly distributed* among the atoms and molecules of a substance, the science of figuring out the implications of randomness is critical for understanding many biological phenomena.

We will begin our study of the implications of microscopic properties and randomness with two phenomena: *kinetic theory* and *diffusion*. Kinetic theory is about understanding thermal phenomena in molecular terms, and diffusion is about what happens when materials are not uniformly distributed. Analyzing both of these using the methods of statistical physics will give us insights into the mechanism of a large class of complex and important phenomena.

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.

1.2 Atomic Structure and Symbolism

The development of modern atomic theory revealed much about the inner structure of atoms. It was learned that an atom contains a very small nucleus composed of positively charged protons and uncharged neutrons, surrounded by a much larger volume of space containing negatively charged electrons. The nucleus contains the majority of an atom's mass because protons and neutrons are much heavier than electrons, whereas electrons occupy almost all of an atom's volume. The diameter of an atom is on the order of 10^{-10} m, whereas the diameter of the nucleus is roughly 10^{-15} m—about 100,000 times smaller. For a perspective about their relative sizes, consider this: If the nucleus were the size of a blueberry, the atom would be about the size of a football stadium (**Figure 1.1**).

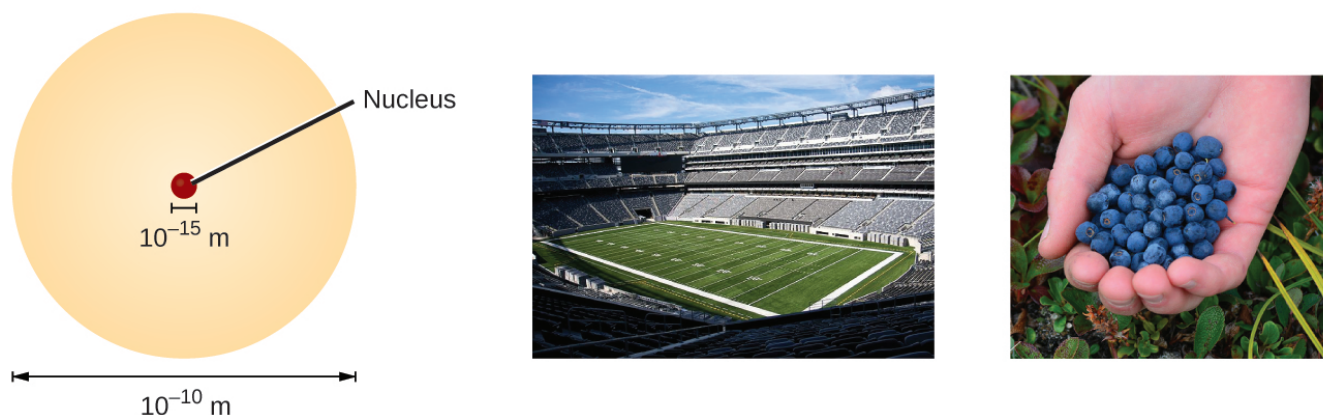


Figure 1.1 If an atom could be expanded to the size of a football stadium, the nucleus would be the size of a single blueberry. (credit middle: modification of work by "babyknight"/Wikimedia Commons; credit right: modification of work by Paxson Woelber)

Atoms—and the protons, neutrons, and electrons that compose them—are extremely small. For example, a carbon atom weighs less than 2×10^{-23} g, and an electron has a charge of less than 2×10^{-19} C (coulomb). When describing the properties of tiny objects such as atoms, we use appropriately small units of measure, such as the **atomic mass unit (amu)** and the **fundamental unit of charge (e)**. The amu was originally defined based on hydrogen, the lightest element, then later in terms of oxygen. Since 1961, it has been defined with regard to the most abundant isotope of carbon, atoms of which are assigned masses of exactly 12 amu. (This isotope is known as “carbon-12” as will be discussed later in this module.) Thus, one amu is exactly $\frac{1}{12}$ of the mass of one carbon-12 atom: $1 \text{ amu} = 1.6605 \times 10^{-24}$ g. (The **Dalton (Da)** and the **unified atomic mass unit (u)** are alternative units that are equivalent to the amu.) The fundamental unit of charge (also called the elementary charge) equals the magnitude of the charge of an electron (e) with $e = 1.602 \times 10^{-19}$ C.

A proton has a mass of 1.0073 amu and a charge of 1+. A neutron is a slightly heavier particle with a mass 1.0087 amu and a charge of zero; as its name suggests, it is neutral. The electron has a charge of 1– and is a much lighter particle with a mass of about 0.00055 amu (it would take about 1800 electrons to equal the mass of one proton). The properties of these fundamental particles are summarized in **Table 1.1**. (An observant student might notice that the sum of an atom’s subatomic particles does not equal the atom’s actual mass: The total mass of six protons, six neutrons, and six electrons is 12.0993 amu, slightly larger than 12.00 amu. This “missing” mass is known as the mass defect, and you will learn about it in the chapter on nuclear chemistry.)

Table 1.1

Properties of Subatomic Particles					
Name	Location	Charge (C)	Unit Charge	Mass (amu)	Mass (g)
electron	outside nucleus	-1.602×10^{-19}	1–	0.00055	0.00091×10^{-24}
proton	nucleus	1.602×10^{-19}	1+	1.00727	1.67262×10^{-24}
neutron	nucleus	0	0	1.00866	1.67493×10^{-24}

The number of protons in the nucleus of an atom is its **atomic number (Z)**. This is the defining trait of an element: Its value determines the identity of the atom. For example, any atom that contains six protons is the element carbon and has the atomic number 6, regardless of how many neutrons or electrons it may have. A neutral atom must contain the same number of positive and negative charges, so the number of protons equals the number of electrons. Therefore, the atomic number also indicates the number of electrons in an atom. The total number of protons and neutrons in an atom is called its **mass number (A)**. The number of neutrons is therefore the difference between the mass number and the atomic number: $A - Z = \text{number of neutrons}$.

$$\begin{aligned}
 \text{atomic number (Z)} &= \text{number of protons} \\
 \text{mass number (A)} &= \text{number of protons} + \text{number of neutrons} \\
 A - Z &= \text{number of neutrons}
 \end{aligned}
 \tag{1.1}$$

Atoms are electrically neutral if they contain the same number of positively charged protons and negatively charged electrons. When the numbers of these subatomic particles are *not* equal, the atom is electrically charged and is called an **ion**. The charge of an atom is defined as follows:

Atomic charge = number of protons – number of electrons

As will be discussed in more detail later in this chapter, atoms (and molecules) typically acquire charge by gaining or losing electrons. An atom that gains one or more electrons will exhibit a negative charge and is called an **anion**. Positively charged

atoms called **cations** are formed when an atom loses one or more electrons. For example, a neutral sodium atom ($Z = 11$) has 11 electrons. If this atom loses one electron, it will become a cation with a $1+$ charge ($11 - 10 = 1+$). A neutral oxygen atom ($Z = 8$) has eight electrons, and if it gains two electrons it will become an anion with a $2-$ charge ($8 - 10 = 2-$).

Example 1.1

Composition of an Atom

Iodine is an essential trace element in our diet; it is needed to produce thyroid hormone. Insufficient iodine in the diet can lead to the development of a goiter, an enlargement of the thyroid gland (**Figure 1.2**).

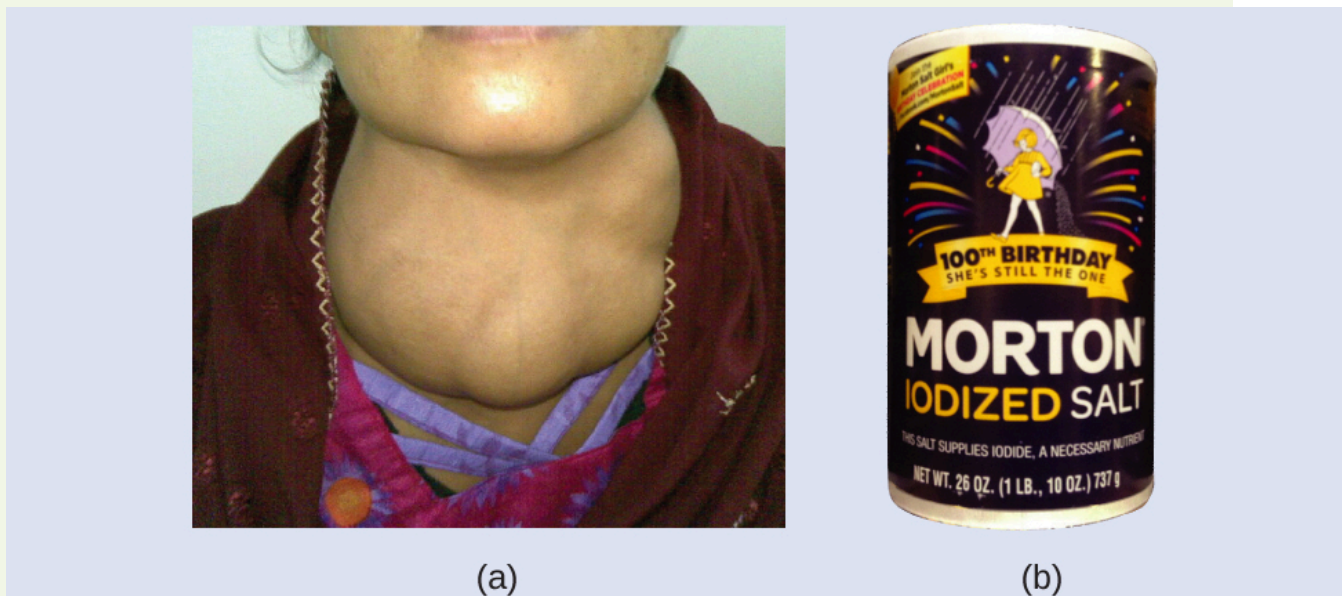


Figure 1.2 (a) Insufficient iodine in the diet can cause an enlargement of the thyroid gland called a goiter. (b) The addition of small amounts of iodine to salt, which prevents the formation of goiters, has helped eliminate this concern in the US where salt consumption is high. (credit a: modification of work by "Almazi"/Wikimedia Commons; credit b: modification of work by Mike Mozart)

The addition of small amounts of iodine to table salt (iodized salt) has essentially eliminated this health concern in the United States, but as much as 40% of the world's population is still at risk of iodine deficiency. The iodine atoms are added as anions, and each has a $1-$ charge and a mass number of 127. Determine the numbers of protons, neutrons, and electrons in one of these iodine anions.

Solution

The atomic number of iodine (53) tells us that a neutral iodine atom contains 53 protons in its nucleus and 53 electrons outside its nucleus. Because the sum of the numbers of protons and neutrons equals the mass number, 127, the number of neutrons is 74 ($127 - 53 = 74$). Since the iodine is added as a $1-$ anion, the number of electrons is 54 [$53 - (1-) = 54$].

Check Your Learning

An ion of platinum has a mass number of 195 and contains 74 electrons. How many protons and neutrons does it contain, and what is its charge?

Answer:

78 protons; 117 neutrons; charge is $4+$

Chemical Symbols

A **chemical symbol** is an abbreviation that we use to indicate an element or an atom of an element. For example, the symbol for mercury is Hg (**Figure 1.3**). We use the same symbol to indicate one atom of mercury (microscopic domain) or to label a container of many atoms of the element mercury (macroscopic domain).



Figure 1.3 The symbol Hg represents the element mercury regardless of the amount; it could represent one atom of mercury or a large amount of mercury.

The symbols for several common elements and their atoms are listed in **Table 1.2**. Some symbols are derived from the common name of the element; others are abbreviations of the name in another language. Most symbols have one or two letters, but three-letter symbols have been used to describe some elements that have atomic numbers greater than 112. To avoid confusion with other notations, only the first letter of a symbol is capitalized. For example, Co is the symbol for the element cobalt, but CO is the notation for the compound carbon monoxide, which contains atoms of the elements carbon (C) and oxygen (O). All known elements and their symbols are in the periodic table in **m51003** (https://legacy.cnx.org/content/m51003/latest/#CNX_Chem_02_05_PerTable1) (also found in **m51209** (<https://legacy.cnx.org/content/m51209/latest/#fs-idp64991424>)).

Table 1.2

Some Common Elements and Their Symbols			
Element	Symbol	Element	Symbol
aluminum	Al	iron	Fe (from <i>ferrum</i>)
bromine	Br	lead	Pb (from <i>plumbum</i>)
calcium	Ca	magnesium	Mg
carbon	C	mercury	Hg (from <i>hydrargyrum</i>)
chlorine	Cl	nitrogen	N
chromium	Cr	oxygen	O
cobalt	Co	potassium	K (from <i>kalium</i>)
copper	Cu (from <i>cuprum</i>)	silicon	Si
fluorine	F	silver	Ag (from <i>argentum</i>)
gold	Au (from <i>aurum</i>)	sodium	Na (from <i>natrium</i>)
helium	He	sulfur	S
hydrogen	H	tin	Sn (from <i>stannum</i>)
iodine	I	zinc	Zn

Traditionally, the discoverer (or discoverers) of a new element names the element. However, until the name is recognized by the International Union of Pure and Applied Chemistry (IUPAC), the recommended name of the new element is based on the Latin word(s) for its atomic number. For example, element 106 was called unnilhexium (Unh), element 107 was called unnilseptium (Uns), and element 108 was called unniloctium (Uno) for several years. These elements are now named after scientists (or occasionally locations); for example, element 106 is now known as *seaborgium* (Sg) in honor of Glenn Seaborg, a Nobel Prize winner who was active in the discovery of several heavy elements.



Visit this [site \(http://openstaxcollege.org/l/16IUPAC\)](http://openstaxcollege.org/l/16IUPAC) to learn more about IUPAC, the International Union of Pure and Applied Chemistry, and explore its periodic table.

Isotopes

The symbol for a specific isotope of any element is written by placing the mass number as a superscript to the left of the element symbol (**Figure 1.4**). The atomic number is sometimes written as a subscript preceding the symbol, but since this number defines the element's identity, as does its symbol, it is often omitted. For example, magnesium exists as a mixture of three isotopes, each with an atomic number of 12 and with mass numbers of 24, 25, and 26, respectively. These isotopes can be identified as ^{24}Mg , ^{25}Mg , and ^{26}Mg . These isotope symbols are read as "element, mass number" and can be symbolized consistent with this reading. For instance, ^{24}Mg is read as "magnesium 24," and can be written as "magnesium-24" or "Mg-24." ^{25}Mg is read as "magnesium 25," and can be written as "magnesium-25" or "Mg-25." All magnesium atoms have 12 protons in their nucleus. They differ only because a ^{24}Mg atom has 12 neutrons in its nucleus, a ^{25}Mg atom has 13 neutrons, and a ^{26}Mg has 14 neutrons.

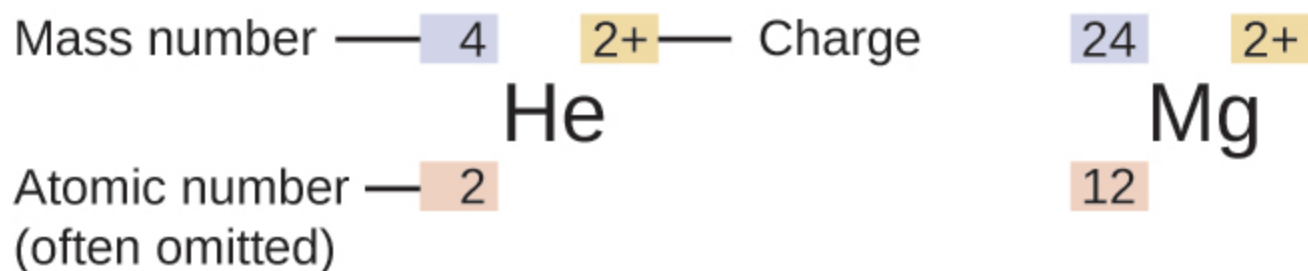


Figure 1.4 The symbol for an atom indicates the element via its usual two-letter symbol, the mass number as a left superscript, the atomic number as a left subscript (sometimes omitted), and the charge as a right superscript.

Information about the naturally occurring isotopes of elements with atomic numbers 1 through 10 is given in **Table 1.3**. Note that in addition to standard names and symbols, the isotopes of hydrogen are often referred to using common names and accompanying symbols. Hydrogen-2, symbolized ^2H , is also called deuterium and sometimes symbolized D. Hydrogen-3, symbolized ^3H , is also called tritium and sometimes symbolized T.

Table 1.3

Nuclear Compositions of Atoms of the Very Light Elements						
Element	Symbol	Atomic Number	Number of Protons	Number of Neutrons	Mass (amu)	% Natural Abundance
hydrogen	${}^1_1\text{H}$ (protium)	1	1	0	1.0078	99.989
	${}^2_1\text{H}$ (deuterium)	1	1	1	2.0141	0.0115
	${}^3_1\text{H}$ (tritium)	1	1	2	3.01605	— (trace)
helium	${}^3_2\text{He}$	2	2	1	3.01603	0.00013
	${}^4_2\text{He}$	2	2	2	4.0026	100
lithium	${}^6_3\text{Li}$	3	3	3	6.0151	7.59
	${}^7_3\text{Li}$	3	3	4	7.0160	92.41
beryllium	${}^9_4\text{Be}$	4	4	5	9.0122	100
boron	${}^{10}_5\text{B}$	5	5	5	10.0129	19.9
	${}^{11}_5\text{B}$	5	5	6	11.0093	80.1
carbon	${}^{12}_6\text{C}$	6	6	6	12.0000	98.89
	${}^{13}_6\text{C}$	6	6	7	13.0034	1.11
	${}^{14}_6\text{C}$	6	6	8	14.0032	— (trace)
nitrogen	${}^{14}_7\text{N}$	7	7	7	14.0031	99.63
	${}^{15}_7\text{N}$	7	7	8	15.0001	0.37
oxygen	${}^{16}_8\text{O}$	8	8	8	15.9949	99.757
	${}^{17}_8\text{O}$	8	8	9	16.9991	0.038
	${}^{18}_8\text{O}$	8	8	10	17.9992	0.205
fluorine	${}^{19}_9\text{F}$	9	9	10	18.9984	100
neon	${}^{20}_{10}\text{Ne}$	10	10	10	19.9924	90.48
	${}^{21}_{10}\text{Ne}$	10	10	11	20.9938	0.27
	${}^{22}_{10}\text{Ne}$	10	10	12	21.9914	9.25



Use this **Build an Atom simulator** (<http://openstaxcollege.org//16PhetAtomBld>) to build atoms of the first 10 elements, see which isotopes exist, check nuclear stability, and gain experience with isotope symbols.

Atomic Mass

Because each proton and each neutron contribute approximately one amu to the mass of an atom, and each electron contributes far less, the **atomic mass** of a single atom is approximately equal to its mass number (a whole number). However, the average masses of atoms of most elements are not whole numbers because most elements exist naturally as mixtures of two or more isotopes.

The mass of an element shown in a periodic table or listed in a table of atomic masses is a weighted, average mass of all the isotopes present in a naturally occurring sample of that element. This is equal to the sum of each individual isotope's mass multiplied by its fractional abundance.

$$\text{average mass} = \sum_i (\text{fractional abundance} \times \text{isotopic mass})_i \quad (1.2)$$

For example, the element boron is composed of two isotopes: About 19.9% of all boron atoms are ^{10}B with a mass of 10.0129 amu, and the remaining 80.1% are ^{11}B with a mass of 11.0093 amu. The average atomic mass for boron is calculated to be:

$$\begin{aligned} \text{boron average mass} &= (0.199 \times 10.0129 \text{ amu}) + (0.801 \times 11.0093 \text{ amu}) \\ &= 1.99 \text{ amu} + 8.82 \text{ amu} \\ &= 10.81 \text{ amu} \end{aligned} \quad (1.3)$$

It is important to understand that no single boron atom weighs exactly 10.8 amu; 10.8 amu is the average mass of all boron atoms, and individual boron atoms weigh either approximately 10 amu or 11 amu.

Example 1.2

Calculation of Average Atomic Mass

A meteorite found in central Indiana contains traces of the noble gas neon picked up from the solar wind during the meteorite's trip through the solar system. Analysis of a sample of the gas showed that it consisted of 91.84% ^{20}Ne (mass 19.9924 amu), 0.47% ^{21}Ne (mass 20.9940 amu), and 7.69% ^{22}Ne (mass 21.9914 amu). What is the average mass of the neon in the solar wind?

Solution

$$\begin{aligned} \text{average mass} &= (0.9184 \times 19.9924 \text{ amu}) + (0.0047 \times 20.9940 \text{ amu}) + (0.0769 \times 21.9914 \text{ amu}) \\ &= (18.36 + 0.099 + 1.69) \text{ amu} \\ &= 20.15 \text{ amu} \end{aligned} \quad (1.4)$$

The average mass of a neon atom in the solar wind is 20.15 amu. (The average mass of a terrestrial neon atom is 20.1796 amu. This result demonstrates that we may find slight differences in the natural abundance of isotopes, depending on their origin.)

Check Your Learning

A sample of magnesium is found to contain 78.70% of ^{24}Mg atoms (mass 23.98 amu), 10.13% of ^{25}Mg atoms (mass 24.99 amu), and 11.17% of ^{26}Mg atoms (mass 25.98 amu). Calculate the average mass of a Mg atom.

Answer:

24.31 amu

We can also do variations of this type of calculation, as shown in the next example.

Example 1.3

Calculation of Percent Abundance

Naturally occurring chlorine consists of ^{35}Cl (mass 34.96885 amu) and ^{37}Cl (mass 36.96590 amu), with an average mass of 35.453 amu. What is the percent composition of Cl in terms of these two isotopes?

Solution

The average mass of chlorine is the fraction that is ^{35}Cl times the mass of ^{35}Cl plus the fraction that is ^{37}Cl times the mass of ^{37}Cl .

$$\text{average mass} = (\text{fraction of } ^{35}\text{Cl} \times \text{mass of } ^{35}\text{Cl}) + (\text{fraction of } ^{37}\text{Cl} \times \text{mass of } ^{37}\text{Cl}) \quad (1.5)$$

If we let x represent the fraction that is ^{35}Cl , then the fraction that is ^{37}Cl is represented by $1.00 - x$.

(The fraction that is ^{35}Cl + the fraction that is ^{37}Cl must add up to 1, so the fraction of ^{37}Cl must equal $1.00 -$ the fraction of ^{35}Cl .)

Substituting this into the average mass equation, we have:

$$\begin{aligned} 35.453 \text{ amu} &= (x \times 34.96885 \text{ amu}) + [(1.00 - x) \times 36.96590 \text{ amu}] \\ 35.453 &= 34.96885x + 36.96590 - 36.96590x \\ 1.99705x &= 1.513 \\ x &= \frac{1.513}{1.99705} = 0.7576 \end{aligned} \quad (1.6)$$

So solving yields: $x = 0.7576$, which means that $1.00 - 0.7576 = 0.2424$. Therefore, chlorine consists of 75.76% ^{35}Cl and 24.24% ^{37}Cl .

Check Your Learning

Naturally occurring copper consists of ^{63}Cu (mass 62.9296 amu) and ^{65}Cu (mass 64.9278 amu), with an average mass of 63.546 amu. What is the percent composition of Cu in terms of these two isotopes?

Answer:

69.15% Cu-63 and 30.85% Cu-65



Visit this [site \(http://openstaxcollege.org/l/16PhetAtomMass\)](http://openstaxcollege.org/l/16PhetAtomMass) to make mixtures of the main isotopes of the first 18 elements, gain experience with average atomic mass, and check naturally occurring isotope ratios using the Isotopes and Atomic Mass simulation.

The occurrence and natural abundances of isotopes can be experimentally determined using an instrument called a mass spectrometer. Mass spectrometry (MS) is widely used in chemistry, forensics, medicine, environmental science, and many other fields to analyze and help identify the substances in a sample of material. In a typical mass spectrometer (**Figure 1.5**), the sample is vaporized and exposed to a high-energy electron beam that causes the sample's atoms (or molecules) to become electrically charged, typically by losing one or more electrons. These cations then pass through a (variable) electric or magnetic field that deflects each cation's path to an extent that depends on both its mass and charge (similar to how the path of a large steel ball bearing rolling past a magnet is deflected to a lesser extent than that of a small steel BB). The ions are detected, and a plot of the relative number of ions generated versus their mass-to-charge ratios (a *mass spectrum*) is made. The height of each vertical feature or peak in a mass spectrum is proportional to the fraction of cations with the specified mass-to-charge ratio. Since its initial use during the development of modern atomic theory, MS has evolved to become a powerful tool for chemical analysis.

in a wide range of applications.

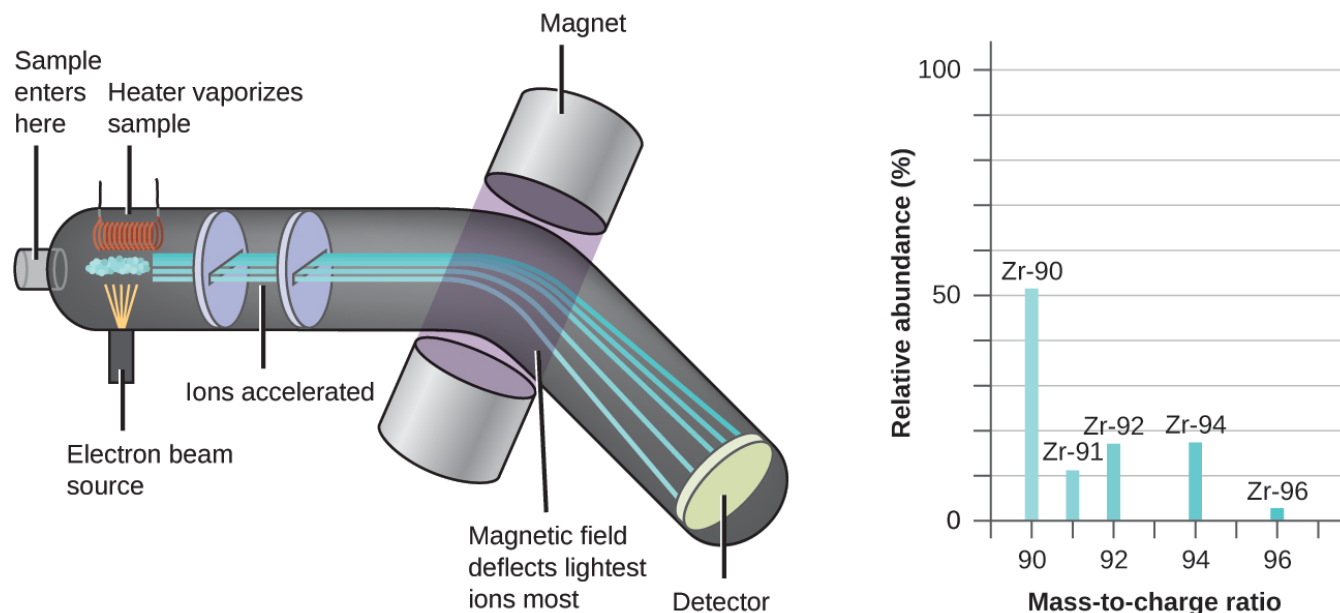


Figure 1.5 Analysis of zirconium in a mass spectrometer produces a mass spectrum with peaks showing the different isotopes of Zr.



See an [animation \(http://openstaxcollege.org//16MassSpec\)](http://openstaxcollege.org//16MassSpec) that explains mass spectrometry. Watch this [video \(http://openstaxcollege.org//16RSCChemistry\)](http://openstaxcollege.org//16RSCChemistry) from the Royal Society for Chemistry for a brief description of the rudiments of mass spectrometry.

Key Concepts and Summary

An atom consists of a small, positively charged nucleus surrounded by electrons. The nucleus contains protons and neutrons; its diameter is about 100,000 times smaller than that of the atom. The mass of one atom is usually expressed in atomic mass units (amu), which is referred to as the atomic mass. An amu is defined as exactly $\frac{1}{12}$ of the mass of a carbon-12 atom and is equal to 1.6605×10^{-24} g.

Protons are relatively heavy particles with a charge of 1+ and a mass of 1.0073 amu. Neutrons are relatively heavy particles with no charge and a mass of 1.0087 amu. Electrons are light particles with a charge of 1- and a mass of 0.00055 amu. The number of protons in the nucleus is called the atomic number (Z) and is the property that defines an atom's elemental identity. The sum of the numbers of protons and neutrons in the nucleus is called the mass number and, expressed in amu, is approximately equal to the mass of the atom. An atom is neutral when it contains equal numbers of electrons and protons.

Isotopes of an element are atoms with the same atomic number but different mass numbers; isotopes of an element, therefore, differ from each other only in the number of neutrons within the nucleus. When a naturally occurring element is composed of several isotopes, the atomic mass of the element represents the average of the masses of the isotopes involved. A chemical symbol identifies the atoms in a substance using symbols, which are one-, two-, or three-letter abbreviations for the atoms.

Key Equations

- average mass = $\sum_i (\text{fractional abundance} \times \text{isotopic mass})_i$

Chemistry End of Chapter Exercises

Exercise 1.1

In what way are isotopes of a given element always different? In what way(s) are they always the same?

Exercise 1.2

Write the symbol for each of the following ions:

- (a) the ion with a 1+ charge, atomic number 55, and mass number 133
- (b) the ion with 54 electrons, 53 protons, and 74 neutrons
- (c) the ion with atomic number 15, mass number 31, and a 3- charge
- (d) the ion with 24 electrons, 30 neutrons, and a 3+ charge

Solution

- (a) $^{133}\text{Cs}^+$; (b) $^{127}\text{I}^-$; (c) $^{31}\text{P}^{3-}$; (d) $^{57}\text{Co}^{3+}$

Exercise 1.3

Write the symbol for each of the following ions:

- (a) the ion with a 3+ charge, 28 electrons, and a mass number of 71
- (b) the ion with 36 electrons, 35 protons, and 45 neutrons
- (c) the ion with 86 electrons, 142 neutrons, and a 4+ charge
- (d) the ion with a 2+ charge, atomic number 38, and mass number 87

Exercise 1.4

Open the **Build an Atom simulation** (<http://openstaxcollege.org/l/16PhetAtomBld>) and click on the Atom icon.

- (a) Pick any one of the first 10 elements that you would like to build and state its symbol.
- (b) Drag protons, neutrons, and electrons onto the atom template to make an atom of your element. State the numbers of protons, neutrons, and electrons in your atom, as well as the net charge and mass number.
- (c) Click on "Net Charge" and "Mass Number," check your answers to (b), and correct, if needed.
- (d) Predict whether your atom will be stable or unstable. State your reasoning.
- (e) Check the "Stable/Unstable" box. Was your answer to (d) correct? If not, first predict what you can do to make a stable atom of your element, and then do it and see if it works. Explain your reasoning.

Solution

- (a) Carbon-12, ^{12}C ; (b) This atom contains six protons and six neutrons. There are six electrons in a neutral ^{12}C atom. The net charge of such a neutral atom is zero, and the mass number is 12. (c) The preceding answers are correct. (d) The atom will be stable since C-12 is a stable isotope of carbon. (e) The preceding answer is correct. Other answers for this exercise are possible if a different element of isotope is chosen.

Exercise 1.5

Open the **Build an Atom simulation** (<http://openstaxcollege.org/l/16PhetAtomBld>)

- (a) Drag protons, neutrons, and electrons onto the atom template to make a neutral atom of Oxygen-16 and give the isotope symbol for this atom.
- (b) Now add two more electrons to make an ion and give the symbol for the ion you have created.

Exercise 1.6

Open the **Build an Atom simulation** (<http://openstaxcollege.org/l/16PhetAtomBld>)

- (a) Drag protons, neutrons, and electrons onto the atom template to make a neutral atom of Lithium-6 and give the isotope symbol for this atom.
- (b) Now remove one electron to make an ion and give the symbol for the ion you have created.

Solution

(a) Lithium-6 contains three protons, three neutrons, and three electrons. The isotope symbol is ${}^6\text{Li}$ or ${}^6_3\text{Li}$. (b) ${}^6\text{Li}^+$ or ${}^6_3\text{Li}^+$

Exercise 1.7

Determine the number of protons, neutrons, and electrons in the following isotopes that are used in medical diagnoses:

- (a) atomic number 9, mass number 18, charge of 1–
- (b) atomic number 43, mass number 99, charge of 7+
- (c) atomic number 53, atomic mass number 131, charge of 1–
- (d) atomic number 81, atomic mass number 201, charge of 1+
- (e) Name the elements in parts (a), (b), (c), and (d).

Exercise 1.8

The following are properties of isotopes of two elements that are essential in our diet. Determine the number of protons, neutrons and electrons in each and name them.

- (a) atomic number 26, mass number 58, charge of 2+
- (b) atomic number 53, mass number 127, charge of 1–

Solution

(a) Iron, 26 protons, 24 electrons, and 32 neutrons; (b) iodine, 53 protons, 54 electrons, and 74 neutrons

Exercise 1.9

Give the number of protons, electrons, and neutrons in neutral atoms of each of the following isotopes:

- (a) ${}^{10}_5\text{B}$
- (b) ${}^{199}_{80}\text{Hg}$
- (c) ${}^{63}_{29}\text{Cu}$
- (d) ${}^{13}_6\text{C}$
- (e) ${}^{77}_{34}\text{Se}$

Exercise 1.10

Give the number of protons, electrons, and neutrons in neutral atoms of each of the following isotopes:

- (a) ${}^7_3\text{Li}$
- (b) ${}^{125}_{52}\text{Te}$
- (c) ${}^{109}_{47}\text{Ag}$
- (d) ${}^{15}_7\text{N}$
- (e) ${}^{31}_{15}\text{P}$

Solution

(a) 3 protons, 3 electrons, 4 neutrons; (b) 52 protons, 52 electrons, 73 neutrons; (c) 47 protons, 47 electrons, 62 neutrons; (d) 7 protons, 7 electrons, 8 neutrons; (e) 15 protons, 15 electrons, 16 neutrons

Exercise 1.11

Click on the [site \(http://openstaxcollege.org/l/16PhetAtomMass\)](http://openstaxcollege.org/l/16PhetAtomMass) and select the “Mix Isotopes” tab, hide the “Percent Composition” and “Average Atomic Mass” boxes, and then select the element boron.

- Write the symbols of the isotopes of boron that are shown as naturally occurring in significant amounts.
- Predict the relative amounts (percentages) of these boron isotopes found in nature. Explain the reasoning behind your choice.
- Add isotopes to the black box to make a mixture that matches your prediction in (b). You may drag isotopes from their bins or click on “More” and then move the sliders to the appropriate amounts.
- Reveal the “Percent Composition” and “Average Atomic Mass” boxes. How well does your mixture match with your prediction? If necessary, adjust the isotope amounts to match your prediction.
- Select “Nature’s” mix of isotopes and compare it to your prediction. How well does your prediction compare with the naturally occurring mixture? Explain. If necessary, adjust your amounts to make them match “Nature’s” amounts as closely as possible.

Exercise 1.12

Repeat **Exercise 1.11** using an element that has three naturally occurring isotopes.

Solution

Let us use neon as an example. Since there are three isotopes, there is no way to be sure to accurately predict the abundances to make the total of 20.18 amu average atomic mass. Let us guess that the abundances are 9% Ne-22, 91% Ne-20, and only a trace of Ne-21. The average mass would be 20.18 amu. Checking the nature’s mix of isotopes shows that the abundances are 90.48% Ne-20, 9.25% Ne-22, and 0.27% Ne-21, so our guessed amounts have to be slightly adjusted.

Exercise 1.13

An element has the following natural abundances and isotopic masses: 90.92% abundance with 19.99 amu, 0.26% abundance with 20.99 amu, and 8.82% abundance with 21.99 amu. Calculate the average atomic mass of this element.

Exercise 1.14

Average atomic masses listed by IUPAC are based on a study of experimental results. Bromine has two isotopes ^{79}Br and ^{81}Br , whose masses (78.9183 and 80.9163 amu) and abundances (50.69% and 49.31%) were determined in earlier experiments. Calculate the average atomic mass of bromine based on these experiments.

Solution

79.904 amu

Exercise 1.15

Variations in average atomic mass may be observed for elements obtained from different sources. Lithium provides an example of this. The isotopic composition of lithium from naturally occurring minerals is 7.5% ^6Li and 92.5% ^7Li , which have masses of 6.01512 amu and 7.01600 amu, respectively. A commercial source of lithium, recycled from a military source, was 3.75% ^6Li (and the rest ^7Li). Calculate the average atomic mass values for each of these two sources.

Exercise 1.16

The average atomic masses of some elements may vary, depending upon the sources of their ores. Naturally occurring boron consists of two isotopes with accurately known masses (^{10}B , 10.0129 amu and ^{11}B , 11.0931 amu). The actual atomic mass of boron can vary from 10.807 to 10.819, depending on whether the mineral source is from Turkey or the United States. Calculate the percent abundances leading to the two values of the average atomic masses of boron from these two countries.

Solution

Turkey source: 26.49% (of 10.0129 amu isotope); US source: 25.37% (of 10.0129 amu isotope)

Exercise 1.17

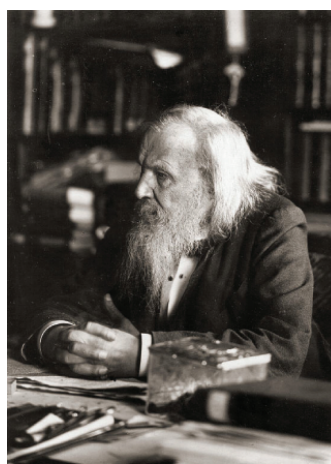
The ^{18}O : ^{16}O abundance ratio in some meteorites is greater than that used to calculate the average atomic mass of oxygen

on earth. Is the average mass of an oxygen atom in these meteorites greater than, less than, or equal to that of a terrestrial oxygen atom?

1.3 The Periodic Table

As early chemists worked to purify ores and discovered more elements, they realized that various elements could be grouped together by their similar chemical behaviors. One such grouping includes lithium (Li), sodium (Na), and potassium (K): These elements all are shiny, conduct heat and electricity well, and have similar chemical properties. A second grouping includes calcium (Ca), strontium (Sr), and barium (Ba), which also are shiny, good conductors of heat and electricity, and have chemical properties in common. However, the specific properties of these two groupings are notably different from each other. For example: Li, Na, and K are much more reactive than are Ca, Sr, and Ba; Li, Na, and K form compounds with oxygen in a ratio of two of their atoms to one oxygen atom, whereas Ca, Sr, and Ba form compounds with one of their atoms to one oxygen atom. Fluorine (F), chlorine (Cl), bromine (Br), and iodine (I) also exhibit similar properties to each other, but these properties are drastically different from those of any of the elements above.

Dimitri Mendeleev in Russia (1869) and Lothar Meyer in Germany (1870) independently recognized that there was a periodic relationship among the properties of the elements known at that time. Both published tables with the elements arranged according to increasing atomic mass. But Mendeleev went one step further than Meyer: He used his table to predict the existence of elements that would have the properties similar to aluminum and silicon, but were yet unknown. The discoveries of gallium (1875) and germanium (1886) provided great support for Mendeleev's work. Although Mendeleev and Meyer had a long dispute over priority, Mendeleev's contributions to the development of the periodic table are now more widely recognized (**Figure 1.6**).



(a)

Reihen	Gruppe I. — R'O	Gruppe II. — R'O	Gruppe III. — R'O ^s	Gruppe IV. RH ⁴ R'O ⁴	Gruppe V. RH ⁵ R'O ⁵	Gruppe VI. RH ⁶ R'O ⁶	Gruppe VII. RH ⁷ R'O ⁷	Gruppe VIII. — R'O ⁴
1	II=1							
2	Li=7	Be=9,4	B=11	C=12	N=14	O=16	F=19	
3	Na=23	Mg=24	Al=27,3	Si=28	P=31	S=32	Cl=35,5	
4	K=39	Ca=40	—=44	Ti=48	V=51	Cr=52	Mn=55	Fe=56, Co=59, Ni=59, Cu=63.
5	(Cu=63)	Zn=65	—=68	—=72	As=75	So=78	Br=80	
6	Rb=86	Sr=87	?Yt=88	Zr=90	Nb=94	Mo=96	—=100	Ru=104, Rh=104, Pd=106, Ag=108.
7	(Ag=108)	Cd=112	In=113	Sn=118	Sb=122	Te=125	J=127	
8	Cs=133	Ba=137	?Di=138	?Co=140	—	—	—	—
9	(—)	—	—	—	—	—	—	—
10	—	—	?Er=178	?La=180	Ta=182	W=184	—	Os=195, Ir=197, Pt=198, Au=199.
11	(Au=199)	Hg=200	Tl=204	Pb=207	Bi=208	—	—	—
12	—	—	—	Th=231	—	U=240	—	—

(b)

Figure 1.6 (a) Dimitri Mendeleev is widely credited with creating (b) the first periodic table of the elements. (credit a: modification of work by Serge Lachinov; credit b: modification of work by "Den fjättrade ankan"/Wikimedia Commons)

By the twentieth century, it became apparent that the periodic relationship involved atomic numbers rather than atomic masses. The modern statement of this relationship, the **periodic law**, is as follows: *the properties of the elements are periodic functions of their atomic numbers*. A modern **periodic table** arranges the elements in increasing order of their atomic numbers and groups atoms with similar properties in the same vertical column (**Figure 1.7**). Each box represents an element and contains its atomic number, symbol, average atomic mass, and (sometimes) name. The elements are arranged in seven horizontal rows, called **periods** or **series**, and 18 vertical columns, called **groups**. Groups are labeled at the top of each column. In the United States, the labels traditionally were numerals with capital letters. However, IUPAC recommends that the numbers 1 through 18 be used, and these labels are more common. For the table to fit on a single page, parts of two of the rows, a total of 14 columns, are usually written below the main body of the table.

Periodic Table of the Elements

Period	Group	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1		1 H 1.008 hydrogen																	2 He 4.003 helium
2		3 Li 6.94 lithium	4 Be 9.012 beryllium											5 B 10.81 boron	6 C 12.01 carbon	7 N 14.01 nitrogen	8 O 16.00 oxygen	9 F 19.00 fluorine	10 Ne 20.18 neon
3		11 Na 22.99 sodium	12 Mg 24.31 magnesium											13 Al 26.98 aluminum	14 Si 28.09 silicon	15 P 30.97 phosphorus	16 S 32.06 sulfur	17 Cl 35.45 chlorine	18 Ar 39.95 argon
4		19 K 39.10 potassium	20 Ca 40.08 calcium	21 Sc 44.96 scandium	22 Ti 47.87 titanium	23 V 50.94 vanadium	24 Cr 52.00 chromium	25 Mn 54.94 manganese	26 Fe 55.85 iron	27 Co 58.93 cobalt	28 Ni 58.69 nickel	29 Cu 63.55 copper	30 Zn 65.38 zinc	31 Ga 69.72 gallium	32 Ge 72.63 germanium	33 As 74.92 arsenic	34 Se 78.97 selenium	35 Br 79.90 bromine	36 Kr 83.80 krypton
5		37 Rb 85.47 rubidium	38 Sr 87.62 strontium	39 Y 88.91 yttrium	40 Zr 91.22 zirconium	41 Nb 92.91 niobium	42 Mo 95.95 molybdenum	43 Tc [97] technetium	44 Ru 101.1 ruthenium	45 Rh 102.9 rhodium	46 Pd 106.4 palladium	47 Ag 107.9 silver	48 Cd 112.4 cadmium	49 In 114.8 indium	50 Sn 118.7 tin	51 Sb 121.8 antimony	52 Te 127.6 tellurium	53 I 126.9 iodine	54 Xe 131.3 xenon
6		55 Cs 132.9 cesium	56 Ba 137.3 barium	57-71 La-Lu * lanthanum series	72 Hf 178.5 hafnium	73 Ta 180.9 tantalum	74 W 183.8 tungsten	75 Re 186.2 rhenium	76 Os 190.2 osmium	77 Ir 192.2 iridium	78 Pt 195.1 platinum	79 Au 197.0 gold	80 Hg 200.6 mercury	81 Tl 204.4 thallium	82 Pb 207.2 lead	83 Bi 209.0 bismuth	84 Po [209] polonium	85 At [210] astatine	86 Rn [222] radon
7		87 Fr [223] francium	88 Ra [226] radium	89-103 Ac-Lr ** actinide series	104 Rf [267] rutherfordium	105 Db [270] dubnium	106 Sg [271] seaborgium	107 Bh [270] bohrium	108 Hs [277] hassium	109 Mt [276] meitnerium	110 Ds [281] darmstadtium	111 Rg [282] roentgenium	112 Cn [285] copernicium	113 Uut [285] ununtrium	114 Fl [289] flerovium	115 Uup [288] ununpentium	116 Lv [293] livermorium	117 Uus [294] ununseptium	118 Uuo [294] ununoctium
					57 La 138.9 lanthanum	58 Ce 140.1 cerium	59 Pr 140.9 praseodymium	60 Nd 144.2 neodymium	61 Pm [145] promethium	62 Sm 150.4 samarium	63 Eu 152.0 europium	64 Gd 157.3 gadolinium	65 Tb 158.9 terbium	66 Dy 162.5 dysprosium	67 Ho 164.9 holmium	68 Er 167.3 erbium	69 Tm 168.9 thulium	70 Yb 173.1 ytterbium	71 Lu 175.0 lutetium
					89 Ac [227] actinium	90 Th 232.0 thorium	91 Pa 231.0 protactinium	92 U 238.0 uranium	93 Np [237] neptunium	94 Pu [244] plutonium	95 Am [243] americium	96 Cm [247] curium	97 Bk [247] berkelium	98 Cf [251] californium	99 Es [252] einsteinium	100 Fm [257] fermium	101 Md [258] mendelevium	102 No [259] nobelium	103 Lr [262] lawrencium

Atomic number → 1

Symbol → **H**

Atomic mass → 1.008

Name → hydrogen

Color Code

Metal	Solid
Metalloid	Liquid
Nonmetal	Gas

Figure 1.7 Elements in the periodic table are organized according to their properties.

1.4 The Mole

Formula Mass

Formula Mass for Covalent Substances

For covalent substances, the formula represents the numbers and types of atoms composing a single molecule of the substance; therefore, the formula mass may be correctly referred to as a molecular mass. Consider chloroform (CHCl_3), a covalent compound once used as a surgical anesthetic and now primarily used in the production of tetrafluoroethylene, the building block for the "anti-stick" polymer, Teflon. The molecular formula of chloroform indicates that a single molecule contains one carbon atom, one hydrogen atom, and three chlorine atoms. The average molecular mass of a chloroform molecule is therefore equal to the sum of the average atomic masses of these atoms. Figure 1.8 outlines the calculations used to derive the molecular mass of chloroform, which is 119.37 amu.

Element	Quantity		Average atomic mass (amu)		Subtotal (amu)
C	1	×	12.01	=	12.01
H	1	×	1.008	=	1.008
Cl	3	×	35.45	=	106.35
Molecular mass					119.37

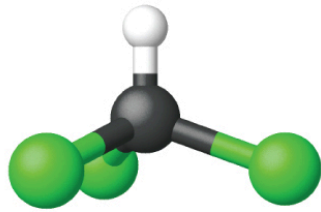


Figure 1.8 The average mass of a chloroform molecule, CHCl_3 , is 119.37 amu, which is the sum of the average atomic masses of each of its constituent atoms. The model shows the molecular structure of chloroform.

Likewise, the molecular mass of an aspirin molecule, $\text{C}_9\text{H}_8\text{O}_4$, is the sum of the atomic masses of nine carbon atoms, eight hydrogen atoms, and four oxygen atoms, which amounts to 180.15 amu (**Figure 1.9**).

Element	Quantity		Average atomic mass (amu)		Subtotal (amu)
C	9	×	12.01	=	108.09
H	8	×	1.008	=	8.064
O	4	×	16.00	=	64.00
Molecular mass					180.15

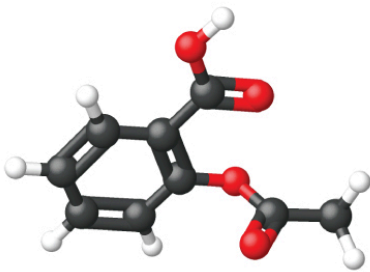


Figure 1.9 The average mass of an aspirin molecule is 180.15 amu. The model shows the molecular structure of aspirin, $\text{C}_9\text{H}_8\text{O}_4$.

Example 1.4

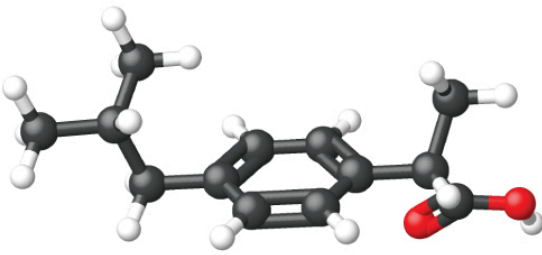
Computing Molecular Mass for a Covalent Compound

Ibuprofen, $\text{C}_{13}\text{H}_{18}\text{O}_2$, is a covalent compound and the active ingredient in several popular nonprescription pain medications, such as Advil and Motrin. What is the molecular mass (amu) for this compound?

Solution

Molecules of this compound are comprised of 13 carbon atoms, 18 hydrogen atoms, and 2 oxygen atoms. Following the approach described above, the average molecular mass for this compound is therefore:

Element	Quantity		Average atomic mass (amu)		Subtotal (amu)
C	13	×	12.01	=	156.13
H	18	×	1.008	=	18.114
O	2	×	16.00	=	32.00
Molecular mass					206.27



Check Your Learning

Acetaminophen, $\text{C}_8\text{H}_9\text{NO}_2$, is a covalent compound and the active ingredient in several popular nonprescription pain medications, such as Tylenol. What is the molecular mass (amu) for this compound?

Answer:

151.16 amu

Formula Mass for Ionic Compounds

Ionic compounds are composed of discrete cations and anions combined in ratios to yield electrically neutral bulk matter. The formula mass for an ionic compound is calculated in the same way as the formula mass for covalent compounds: by summing the average atomic masses of all the atoms in the compound's formula. Keep in mind, however, that the formula for an ionic compound does not represent the composition of a discrete molecule, so it may not correctly be referred to as the "molecular mass."

As an example, consider sodium chloride, NaCl, the chemical name for common table salt. Sodium chloride is an ionic compound composed of sodium cations, Na^+ , and chloride anions, Cl^- , combined in a 1:1 ratio. The formula mass for this compound is computed as 58.44 amu (see **Figure 1.10**).

Element	Quantity		Average atomic mass (amu)		Subtotal
Na	1	×	22.99	=	22.99
Cl	1	×	35.45	=	35.45
Formula mass					58.44

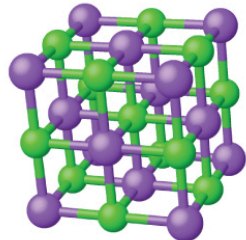


Figure 1.10 Table salt, NaCl, contains an array of sodium and chloride ions combined in a 1:1 ratio. Its formula mass is 58.44 amu.

Note that the average masses of neutral sodium and chlorine atoms were used in this computation, rather than the masses for sodium cations and chloride anions. This approach is perfectly acceptable when computing the formula mass of an ionic compound. Even though a sodium cation has a slightly smaller mass than a sodium atom (since it is missing an electron), this difference will be offset by the fact that a chloride anion is slightly more massive than a chloride atom (due to the extra electron). Moreover, the mass of an electron is negligibly small with respect to the mass of a typical atom. Even when calculating the mass of an isolated ion, the missing or additional electrons can generally be ignored, since their contribution to the overall mass is negligible, reflected only in the nonsignificant digits that will be lost when the computed mass is properly rounded. The few exceptions to this guideline are very light ions derived from elements with precisely known atomic masses.

Example 1.5

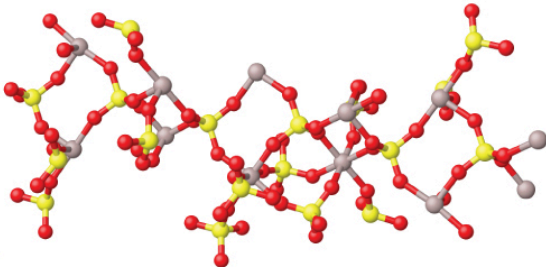
Computing Formula Mass for an Ionic Compound

Aluminum sulfate, $\text{Al}_2(\text{SO}_4)_3$, is an ionic compound that is used in the manufacture of paper and in various water purification processes. What is the formula mass (amu) of this compound?

Solution

The formula for this compound indicates it contains Al^{3+} and SO_4^{2-} ions combined in a 2:3 ratio. For purposes of computing a formula mass, it is helpful to rewrite the formula in the simpler format, $\text{Al}_2\text{S}_3\text{O}_{12}$. Following the approach outlined above, the formula mass for this compound is calculated as follows:

Element	Quantity		Average atomic mass (amu)		Subtotal (amu)
Al	2	×	26.98	=	53.96
S	3	×	32.06	=	96.18
O	12	×	16.00	=	192.00
Molecular mass					342.14



Check Your Learning

Calcium phosphate, $\text{Ca}_3(\text{PO}_4)_2$, is an ionic compound and a common anti-caking agent added to food products. What is the formula mass (amu) of calcium phosphate?

Answer:

310.18 amu

The Mole

The identity of a substance is defined not only by the types of atoms or ions it contains, but by the quantity of each type of atom or ion. For example, water, H_2O , and hydrogen peroxide, H_2O_2 , are alike in that their respective molecules are composed of hydrogen and oxygen atoms. However, because a hydrogen peroxide molecule contains two oxygen atoms, as opposed to the water molecule, which has only one, the two substances exhibit very different properties. Today, we possess sophisticated instruments that allow the direct measurement of these defining microscopic traits; however, the same traits were originally derived from the measurement of macroscopic properties (the masses and volumes of bulk quantities of matter) using relatively simple tools (balances and volumetric glassware). This experimental approach required the introduction of a new unit for amount of substances, the *mole*, which remains indispensable in modern chemical science.

The mole is an amount unit similar to familiar units like pair, dozen, gross, etc. It provides a specific measure of *the number* of atoms or molecules in a bulk sample of matter. A **mole** is defined as the amount of substance containing the same number of discrete entities (such as atoms, molecules, and ions) as the number of atoms in a sample of pure ^{12}C weighing exactly 12 g. One Latin connotation for the word “mole” is “large mass” or “bulk,” which is consistent with its use as the name for this unit. The mole provides a link between an easily measured macroscopic property, bulk mass, and an extremely important fundamental property, number of atoms, molecules, and so forth.

The number of entities composing a mole has been experimentally determined to be $6.02214179 \times 10^{23}$, a fundamental constant named **Avogadro's number (N_A)** or the Avogadro constant in honor of Italian scientist Amedeo Avogadro. This constant is properly reported with an explicit unit of “per mole,” a conveniently rounded version being $6.022 \times 10^{23}/\text{mol}$.

Consistent with its definition as an amount unit, 1 mole of any element contains the same number of atoms as 1 mole of any other element. The masses of 1 mole of different elements, however, are different, since the masses of the individual atoms are drastically different. The **molar mass** of an element (or compound) is the mass in grams of 1 mole of that substance, a property expressed in units of grams per mole (g/mol) (see **Figure 1.11**).

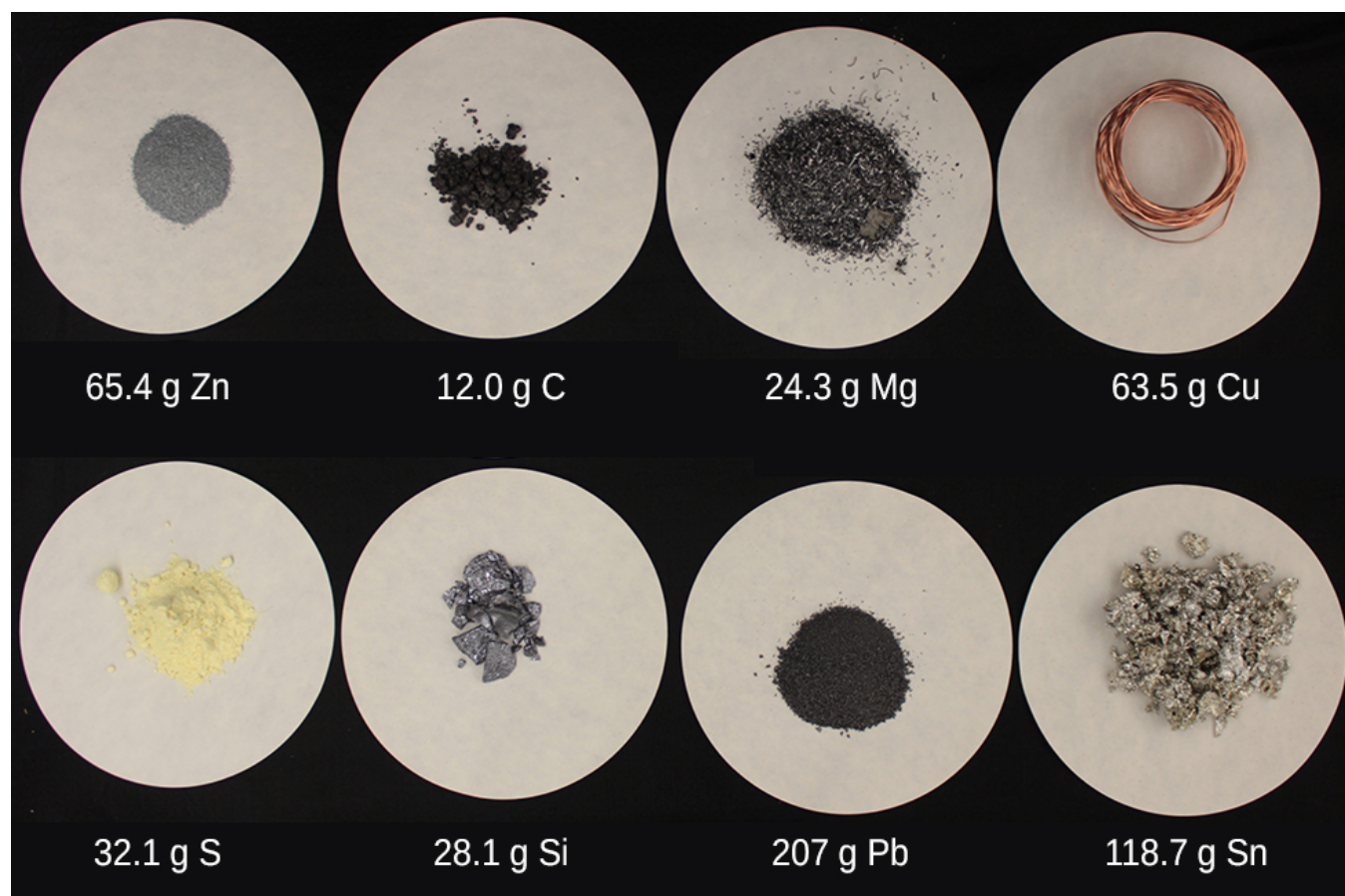


Figure 1.11 Each sample contains 6.022×10^{23} atoms—1.00 mol of atoms. From left to right (top row): 65.4 g zinc, 12.0 g carbon, 24.3 g magnesium, and 63.5 g copper. From left to right (bottom row): 32.1 g sulfur, 28.1 g silicon, 207 g lead, and 118.7 g tin. (credit: modification of work by Mark Ott)

Because the definitions of both the mole and the atomic mass unit are based on the same reference substance, ^{12}C , the molar mass of any substance is numerically equivalent to its atomic or formula weight in amu. Per the amu definition, a single ^{12}C atom weighs 12 amu (its atomic mass is 12 amu). According to the definition of the mole, 12 g of ^{12}C contains 1 mole of ^{12}C atoms (its molar mass is 12 g/mol). This relationship holds for all elements, since their atomic masses are measured relative to that of the amu-reference substance, ^{12}C . Extending this principle, the molar mass of a compound in grams is likewise numerically equivalent to its formula mass in amu (**Figure 1.12**).



Figure 1.12 Each sample contains 6.02×10^{23} molecules or formula units—1.00 mol of the compound or element. Clock-wise from the upper left: 130.2 g of $\text{C}_8\text{H}_{17}\text{OH}$ (1-octanol, formula mass 130.2 amu), 454.4 g of HgI_2 (mercury(II) iodide, formula mass 454.4 amu), 32.0 g of CH_3OH (methanol, formula mass 32.0 amu) and 256.5 g of S_8 (sulfur, formula mass 256.5 amu). (credit: Sahar Atwa)

Table 1.4

Element	Average Atomic Mass (amu)	Molar Mass (g/mol)	Atoms/Mole
C	12.01	12.01	6.022×10^{23}
H	1.008	1.008	6.022×10^{23}
O	16.00	16.00	6.022×10^{23}
Na	22.99	22.99	6.022×10^{23}
Cl	35.45	33.45	6.022×10^{23}

While atomic mass and molar mass are numerically equivalent, keep in mind that they are vastly different in terms of scale, as represented by the vast difference in the magnitudes of their respective units (amu versus g). To appreciate the enormity of the mole, consider a small drop of water weighing about 0.03 g (see **Figure 1.13**). Although this represents just a tiny fraction of 1 mole of water (~18 g), it contains more water molecules than can be clearly imagined. If the molecules were distributed equally among the roughly seven billion people on earth, each person would receive more than 100 billion molecules.

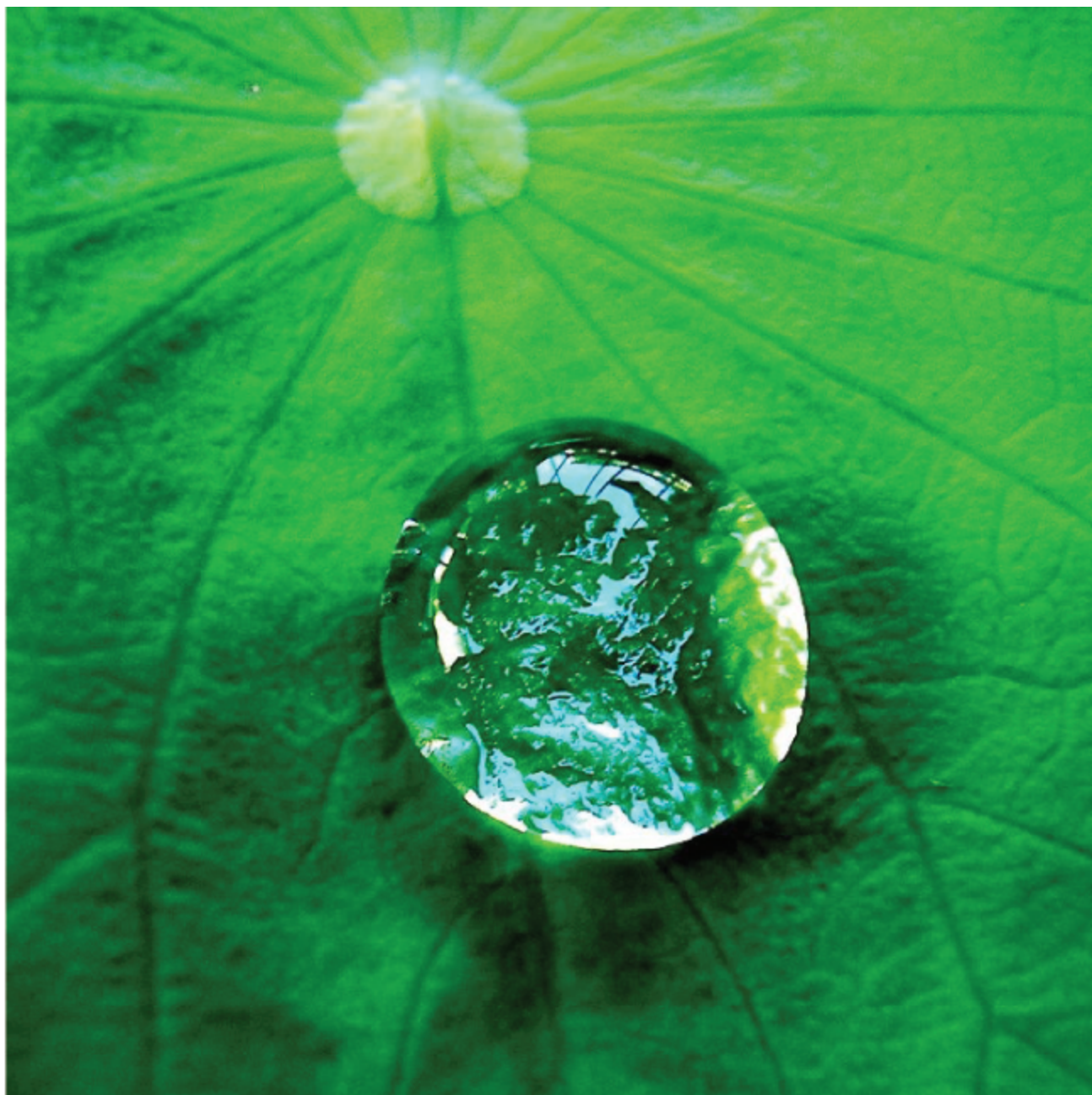


Figure 1.13 The number of molecules in a single droplet of water is roughly 100 billion times greater than the number of people on earth. (credit: "tanakawho"/Wikimedia commons)



The mole is used in chemistry to represent 6.022×10^{23} of something, but it can be difficult to conceptualize such a large number. Watch this [video \(http://openstaxcollege.org/l/16molevideo\)](http://openstaxcollege.org/l/16molevideo) and then complete the “Think” questions that follow. Explore more about the mole by reviewing the information under “Dig Deeper.”

The relationships between formula mass, the mole, and Avogadro's number can be applied to compute various quantities that describe the composition of substances and compounds. For example, if we know the mass and chemical composition of a substance, we can determine the number of moles and calculate number of atoms or molecules in the sample. Likewise, if we know the number of moles of a substance, we can derive the number of atoms or molecules and calculate the substance's mass.

Example 1.6

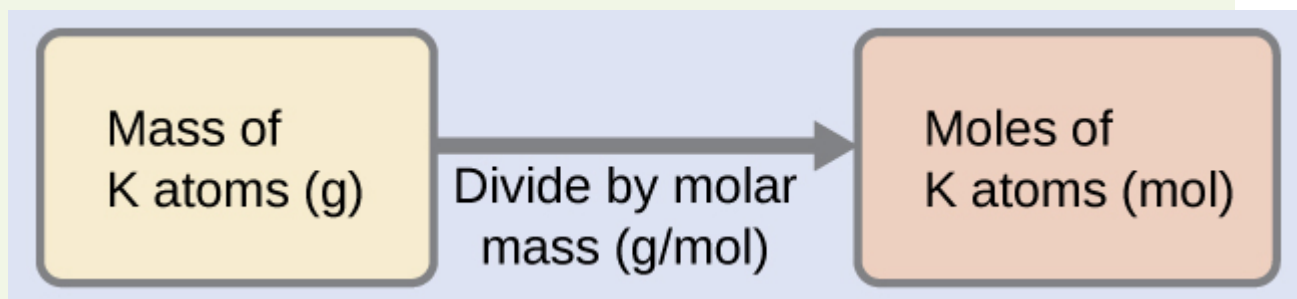
Deriving Moles from Grams for an Element

According to nutritional guidelines from the US Department of Agriculture, the estimated average requirement for dietary potassium is 4.7 g. What is the estimated average requirement of potassium in moles?

Solution

The mass of K is provided, and the corresponding amount of K in moles is requested. Referring to the periodic table, the atomic mass of K is 39.10 amu, and so its molar mass is 39.10 g/mol. The given mass of K (4.7 g) is a bit more than one-tenth the molar mass (39.10 g), so a reasonable "ballpark" estimate of the number of moles would be slightly greater than 0.1 mol.

The molar amount of a substance may be calculated by dividing its mass (g) by its molar mass (g/mol):



The factor-label method supports this mathematical approach since the unit "g" cancels and the answer has units of "mol:"

$$4.7 \text{ g K} \left(\frac{\text{mol K}}{39.10 \text{ g}} \right) = 0.12 \text{ mol K} \quad (1.7)$$

The calculated magnitude (0.12 mol K) is consistent with our ballpark expectation, since it is a bit greater than 0.1 mol.

Check Your Learning

Beryllium is a light metal used to fabricate transparent X-ray windows for medical imaging instruments. How many moles of Be are in a thin-foil window weighing 3.24 g?

Answer:

0.360 mol

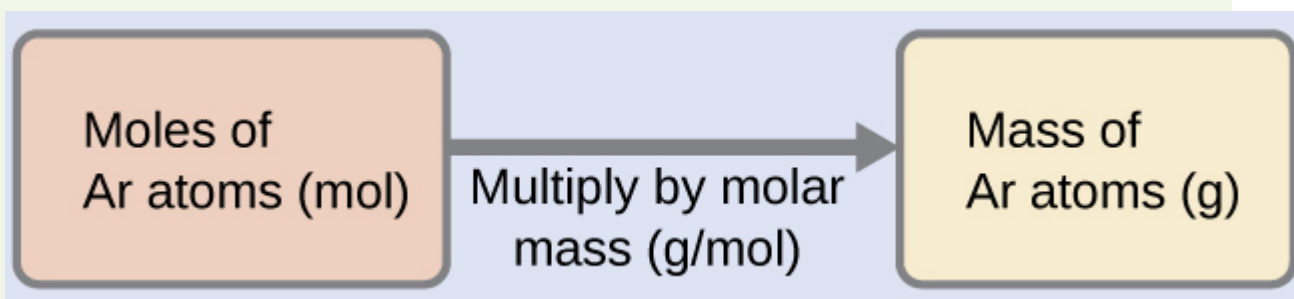
Example 1.7

Deriving Grams from Moles for an Element

A liter of air contains 9.2×10^{-4} mol argon. What is the mass of Ar in a liter of air?

Solution

The molar amount of Ar is provided and must be used to derive the corresponding mass in grams. Since the amount of Ar is less than 1 mole, the mass will be less than the mass of 1 mole of Ar, approximately 40 g. The molar amount in question is approximately one-one thousandth ($\sim 10^{-3}$) of a mole, and so the corresponding mass should be roughly one-one thousandth of the molar mass (~ 0.04 g):



In this case, logic dictates (and the factor-label method supports) multiplying the provided amount (mol) by the molar mass (g/mol):

$$9.2 \times 10^{-4} \text{ mol Ar} \left(\frac{39.95 \text{ g}}{\text{mol Ar}} \right) = 0.037 \text{ g Ar} \quad (1.8)$$

The result is in agreement with our expectations, around 0.04 g Ar.

Check Your Learning

What is the mass of 2.561 mol of gold?

Answer:

504.4 g

Example 1.8

Deriving Number of Atoms from Mass for an Element

Copper is commonly used to fabricate electrical wire (**Figure 1.14**). How many copper atoms are in 5.00 g of copper wire?

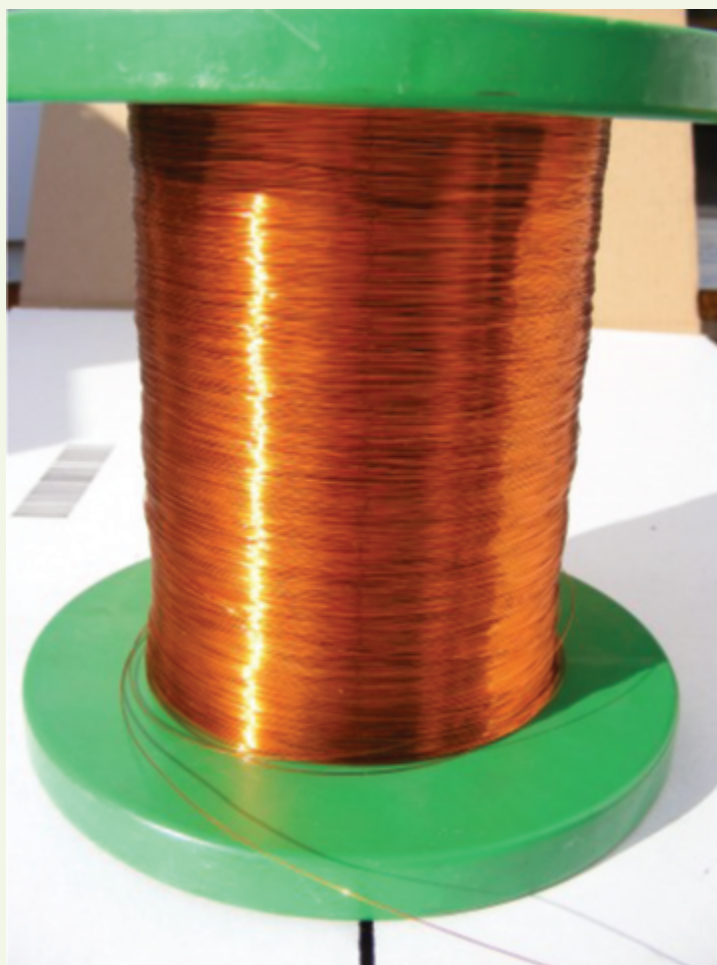
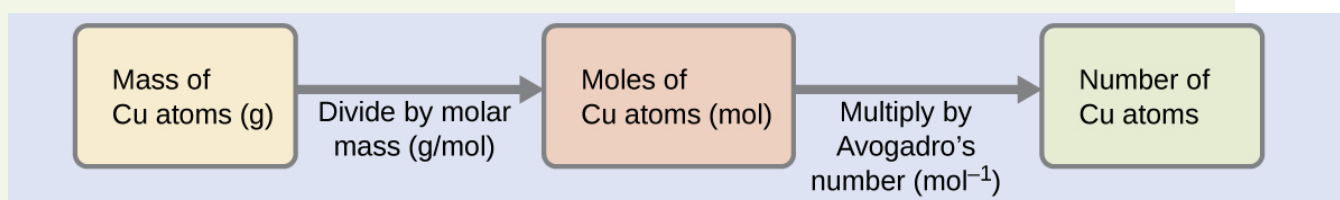


Figure 1.14 Copper wire is composed of many, many atoms of Cu. (credit: Emilian Robert Vicol)

Solution

The number of Cu atoms in the wire may be conveniently derived from its mass by a two-step computation: first calculating the molar amount of Cu, and then using Avogadro's number (N_A) to convert this molar amount to number of Cu atoms:



Considering that the provided sample mass (5.00 g) is a little less than one-tenth the mass of 1 mole of Cu (~64 g), a reasonable estimate for the number of atoms in the sample would be on the order of one-tenth N_A , or approximately 10^{22} Cu atoms. Carrying out the two-step computation yields:

$$5.00 \text{ g Cu} \left(\frac{\text{mol Cu}}{63.55 \text{ g}} \right) \left(\frac{6.022 \times 10^{23} \text{ atoms}}{\text{mol}} \right) = 4.74 \times 10^{22} \text{ atoms of copper} \quad (1.9)$$

The factor-label method yields the desired cancellation of units, and the computed result is on the order of 10^{22} as expected.

Check Your Learning

A prospector panning for gold in a river collects 15.00 g of pure gold. How many Au atoms are in this quantity of gold?

Answer:

$$4.586 \times 10^{22} \text{ Au atoms}$$

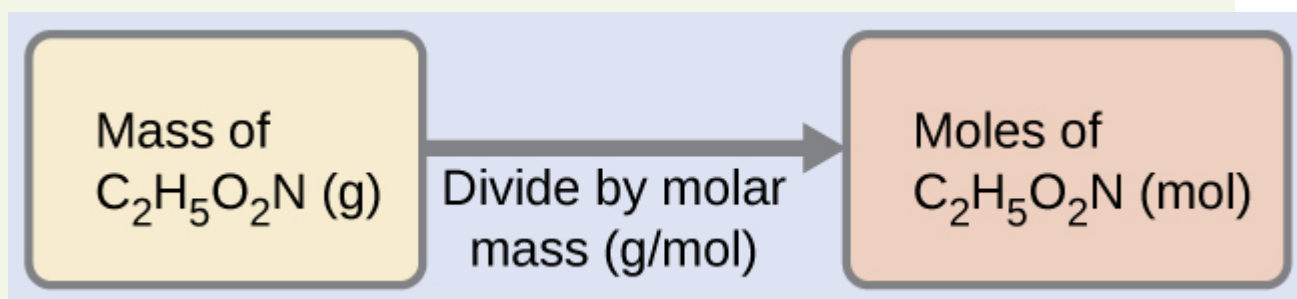
Example 1.9

Deriving Moles from Grams for a Compound

Our bodies synthesize protein from amino acids. One of these amino acids is glycine, which has the molecular formula $\text{C}_2\text{H}_5\text{O}_2\text{N}$. How many moles of glycine molecules are contained in 28.35 g of glycine?

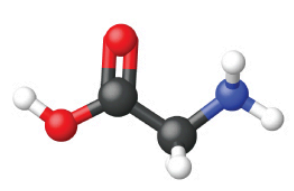
Solution

We can derive the number of moles of a compound from its mass following the same procedure we used for an element in **Example 1.6**:



The molar mass of glycine is required for this calculation, and it is computed in the same fashion as its molecular mass. One mole of glycine, $\text{C}_2\text{H}_5\text{O}_2\text{N}$, contains 2 moles of carbon, 5 moles of hydrogen, 2 moles of oxygen, and 1 mole of nitrogen:

Element	Quantity (mol element/ mol compound)		Molar mass (g/mol element)		Subtotal (g/mol compound)
C	2	×	12.01	=	24.02
H	5	×	1.008	=	5.040
O	2	×	16.00	=	32.00
N	1	×	14.007	=	14.007
Molecular mass (g/mol compound)					75.07



The provided mass of glycine (~28 g) is a bit more than one-third the molar mass (~75 g/mol), so we would expect the computed result to be a bit greater than one-third of a mole (~0.33 mol). Dividing the compound's mass by its molar mass yields:

$$28.35 \text{ g glycine} \left(\frac{\text{mol glycine}}{75.07 \text{ g}} \right) = 0.378 \text{ mol glycine} \quad (1.10)$$

This result is consistent with our rough estimate.

Check Your Learning

How many moles of sucrose, $\text{C}_{12}\text{H}_{22}\text{O}_{11}$, are in a 25-g sample of sucrose?

Answer:

0.073 mol

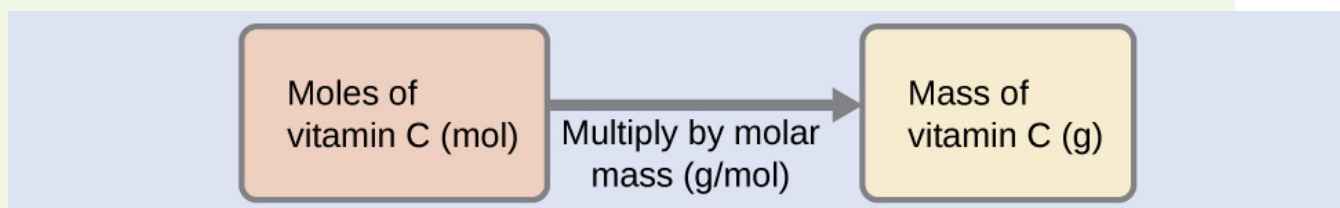
Example 1.10

Deriving Grams from Moles for a Compound

Vitamin C is a covalent compound with the molecular formula $C_6H_8O_6$. The recommended daily dietary allowance of vitamin C for children aged 4–8 years is 1.42×10^{-4} mol. What is the mass of this allowance in grams?

Solution

As for elements, the mass of a compound can be derived from its molar amount as shown:



The molar mass for this compound is computed to be 176.124 g/mol. The given number of moles is a very small fraction of a mole ($\sim 10^{-4}$ or one-ten thousandth); therefore, we would expect the corresponding mass to be about one-ten thousandth of the molar mass (~ 0.02 g). Performing the calculation, we get:

$$1.42 \times 10^{-4} \text{ mol vitamin C} \left(\frac{176.124 \text{ g}}{\text{mol vitamin C}} \right) = 0.0250 \text{ g vitamin C} \quad (1.11)$$

This is consistent with the anticipated result.

Check Your Learning

What is the mass of 0.443 mol of hydrazine, N_2H_4 ?

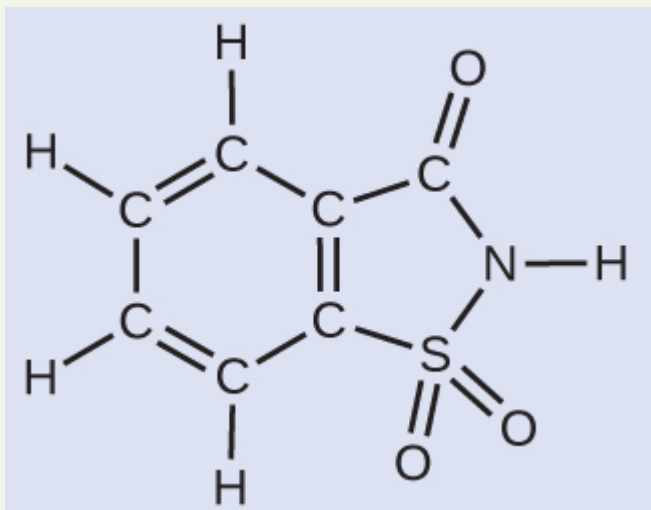
Answer:

14.2 g

Example 1.11

Deriving the Number of Atoms and Molecules from the Mass of a Compound

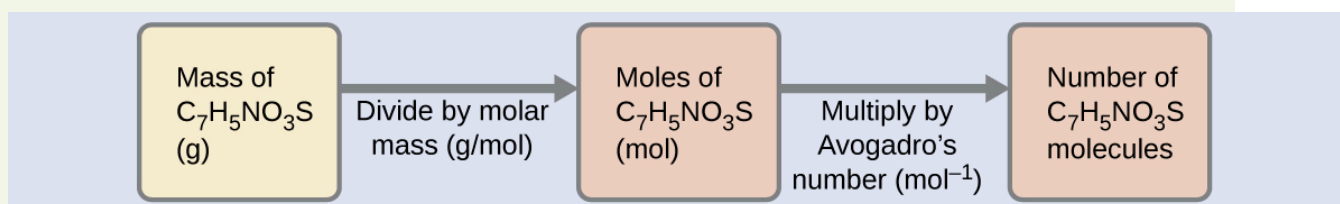
A packet of an artificial sweetener contains 40.0 mg of saccharin ($C_7H_5NO_3S$), which has the structural formula:



Given that saccharin has a molar mass of 183.18 g/mol, how many saccharin molecules are in a 40.0-mg (0.0400-g) sample of saccharin? How many carbon atoms are in the same sample?

Solution

The number of molecules in a given mass of compound is computed by first deriving the number of moles, as demonstrated in **Example 1.9**, and then multiplying by Avogadro's number:



Using the provided mass and molar mass for saccharin yields:

$$0.0400 \text{ g C}_7\text{H}_5\text{NO}_3\text{S} \left(\frac{1 \text{ mol C}_7\text{H}_5\text{NO}_3\text{S}}{183.18 \text{ g C}_7\text{H}_5\text{NO}_3\text{S}} \right) \left(\frac{6.022 \times 10^{23} \text{ C}_7\text{H}_5\text{NO}_3\text{S molecules}}{1 \text{ mol C}_7\text{H}_5\text{NO}_3\text{S}} \right) \quad (1.12)$$

$$= 1.31 \times 10^{20} \text{ C}_7\text{H}_5\text{NO}_3\text{S molecules}$$

The compound's formula shows that each molecule contains seven carbon atoms, and so the number of C atoms in the provided sample is:

$$1.31 \times 10^{20} \text{ C}_7\text{H}_5\text{NO}_3\text{S molecules} \left(\frac{7 \text{ C atoms}}{1 \text{ C}_7\text{H}_5\text{NO}_3\text{S molecule}} \right) = 9.17 \times 10^{20} \text{ C atoms} \quad (1.13)$$

Check Your Learning

How many C_4H_{10} molecules are contained in 9.213 g of this compound? How many hydrogen atoms?

Answer:

$$9.545 \times 10^{22} \text{ molecules C}_4\text{H}_{10}; 9.545 \times 10^{23} \text{ atoms H}$$

Counting Neurotransmitter Molecules in the Brain

The brain is the control center of the central nervous system (**Figure 1.15**). It sends and receives signals to and from muscles and other internal organs to monitor and control their functions; it processes stimuli detected by sensory organs to guide interactions with the external world; and it houses the complex physiological processes that give rise to our intellect and emotions. The broad field of neuroscience spans all aspects of the structure and function of the central nervous system, including research on the anatomy and physiology of the brain. Great progress has been made in brain research over the past few decades, and the BRAIN Initiative, a federal initiative announced in 2013, aims to accelerate and capitalize on

these advances through the concerted efforts of various industrial, academic, and government agencies (more details available at www.whitehouse.gov/share/brain-initiative).

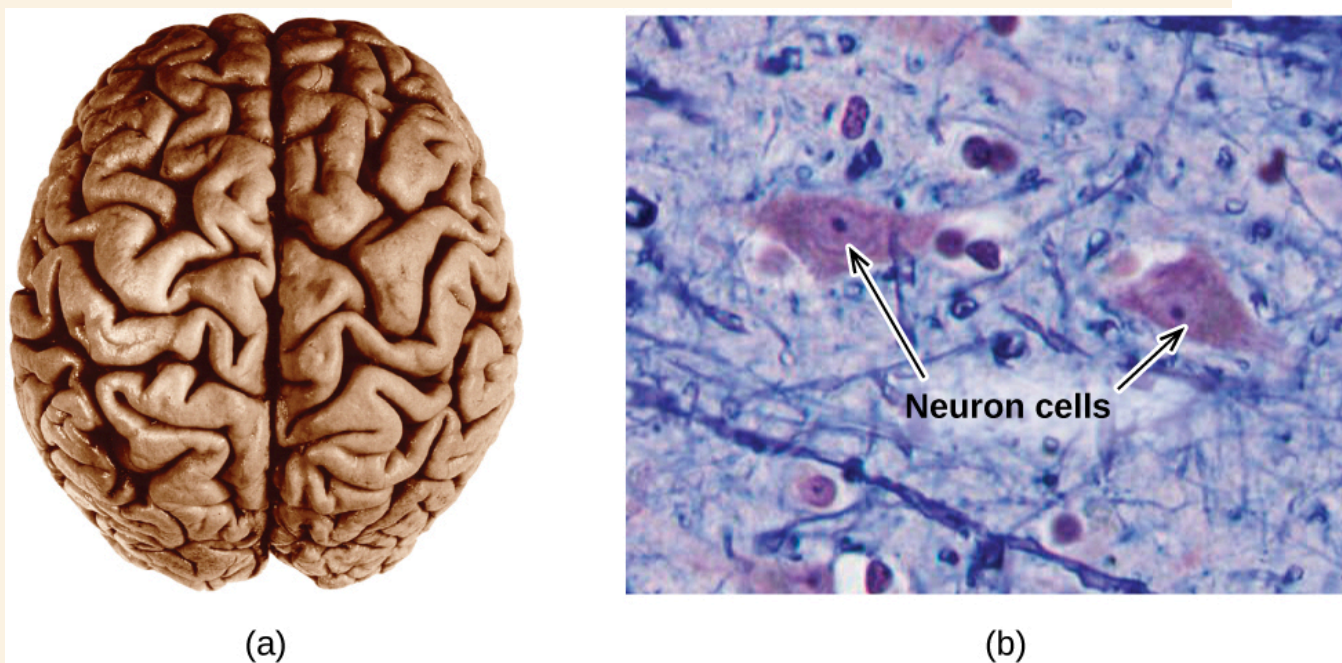


Figure 1.15 (a) A typical human brain weighs about 1.5 kg and occupies a volume of roughly 1.1 L. (b) Information is transmitted in brain tissue and throughout the central nervous system by specialized cells called neurons (micrograph shows cells at 1600× magnification).

Specialized cells called neurons transmit information between different parts of the central nervous system by way of electrical and chemical signals. Chemical signaling occurs at the interface between different neurons when one of the cells releases molecules (called neurotransmitters) that diffuse across the small gap between the cells (called the synapse) and bind to the surface of the other cell. These neurotransmitter molecules are stored in small intracellular structures called vesicles that fuse to the cell wall and then break open to release their contents when the neuron is appropriately stimulated. This process is called exocytosis (see **Figure 1.16**). One neurotransmitter that has been very extensively studied is dopamine, $C_8H_{11}NO_2$. Dopamine is involved in various neurological processes that impact a wide variety of human behaviors. Dysfunctions in the dopamine systems of the brain underlie serious neurological diseases such as Parkinson's and schizophrenia.

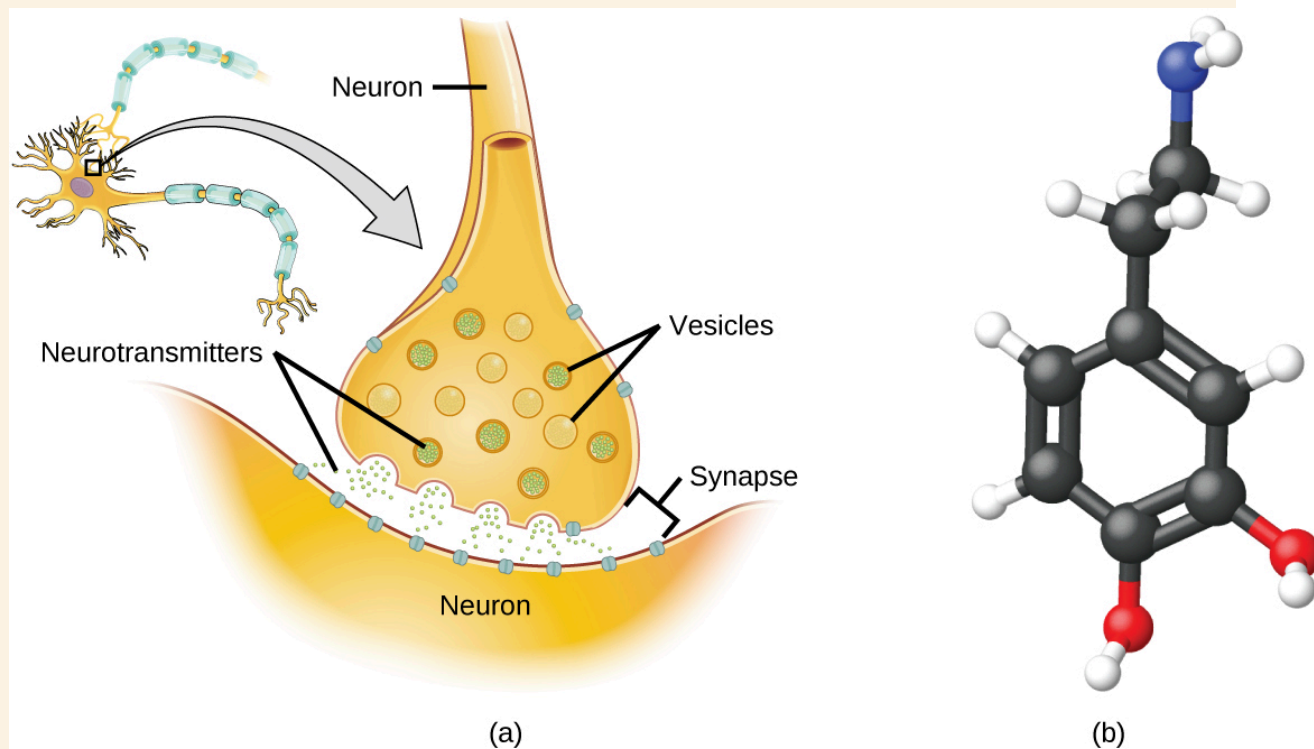


Figure 1.16 (a) Chemical signals are transmitted from neurons to other cells by the release of neurotransmitter molecules into the small gaps (synapses) between the cells. (b) Dopamine, $C_8H_{11}NO_2$, is a neurotransmitter involved in a number of neurological processes.

One important aspect of the complex processes related to dopamine signaling is the number of neurotransmitter molecules released during exocytosis. Since this number is a central factor in determining neurological response (and subsequent human thought and action), it is important to know how this number changes with certain controlled stimulations, such as the administration of drugs. It is also important to understand the mechanism responsible for any changes in the number of neurotransmitter molecules released—for example, some dysfunction in exocytosis, a change in the number of vesicles in the neuron, or a change in the number of neurotransmitter molecules in each vesicle.

Significant progress has been made recently in directly measuring the number of dopamine molecules stored in individual vesicles and the amount actually released when the vesicle undergoes exocytosis. Using miniaturized probes that can selectively detect dopamine molecules in very small amounts, scientists have determined that the vesicles of a certain type of mouse brain neuron contain an average of 30,000 dopamine molecules per vesicle (about 5×10^{-20} mol or 50 zmol). Analysis of these neurons from mice subjected to various drug therapies shows significant changes in the average number of dopamine molecules contained in individual vesicles, increasing or decreasing by up to three-fold, depending on the specific drug used. These studies also indicate that not all of the dopamine in a given vesicle is released during exocytosis, suggesting that it may be possible to regulate the fraction released using pharmaceutical therapies.^[1]

Key Concepts and Summary

The formula mass of a substance is the sum of the average atomic masses of each atom represented in the chemical formula and is expressed in atomic mass units. The formula mass of a covalent compound is also called the molecular mass. A convenient amount unit for expressing very large numbers of atoms or molecules is the mole. Experimental measurements have determined the number of entities composing 1 mole of substance to be 6.022×10^{23} , a quantity called Avogadro's number. The mass in grams of 1 mole of substance is its molar mass. Due to the use of the same reference substance in defining the atomic mass unit and the mole, the formula mass (amu) and molar mass (g/mol) for any substance are numerically equivalent (for example, one H_2O molecule weighs approximately 18 amu and 1 mole of H_2O molecules weighs approximately 18 g).

1. Omiattek, Donna M., Amanda J. Bressler, Ann-Sofie Cans, Anne M. Andrews, Michael L. Heien, and Andrew G. Ewing. "The Real Catecholamine Content of Secretory Vesicles in the CNS Revealed by Electrochemical Cytometry." *Scientific Report* 3 (2013): 1447, accessed January 14, 2015, doi:10.1038/srep01447.

Chemistry End of Chapter Exercises

Exercise 1.18

What is the total mass (amu) of carbon in each of the following molecules?

- (a) CH_4
- (b) CHCl_3
- (c) $\text{C}_{12}\text{H}_{10}\text{O}_6$
- (d) $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3$

Solution

(a) 12.01 amu; (b) 12.01 amu; (c) 144.12 amu; (d) 60.05 amu

Exercise 1.19

What is the total mass of hydrogen in each of the molecules?

- (a) CH_4
- (b) CHCl_3
- (c) $\text{C}_{12}\text{H}_{10}\text{O}_6$
- (d) $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3$

Exercise 1.20

Calculate the molecular or formula mass of each of the following:

- (a) P_4
- (b) H_2O
- (c) $\text{Ca}(\text{NO}_3)_2$
- (d) $\text{CH}_3\text{CO}_2\text{H}$ (acetic acid)
- (e) $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ (sucrose, cane sugar).

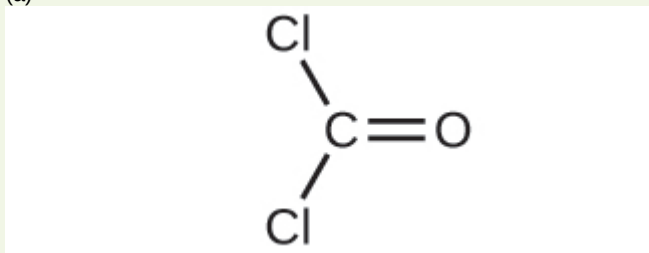
Solution

(a) 123.896 amu; (b) 18.015 amu; (c) 164.086 amu; (d) 60.052 amu; (e) 342.297 amu

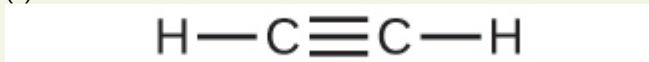
Exercise 1.21

Determine the molecular mass of the following compounds:

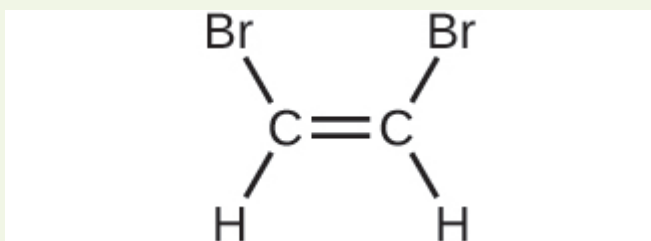
(a)



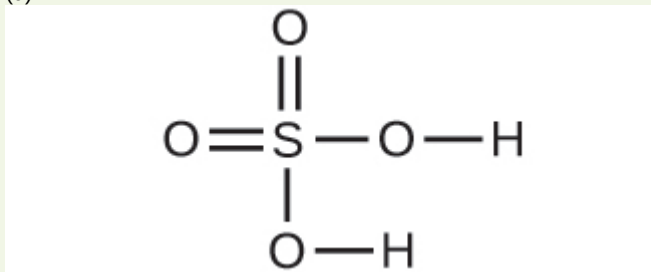
(b)



(c)

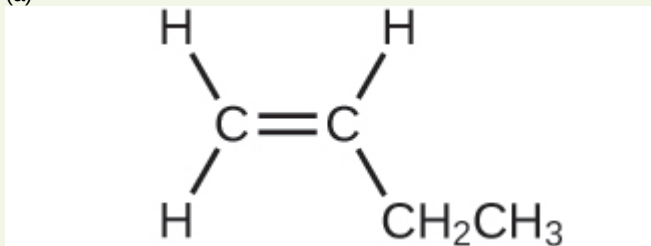


(d)

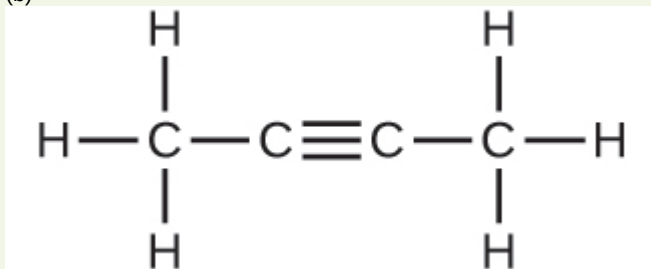
**Exercise 1.22**

Determine the molecular mass of the following compounds:

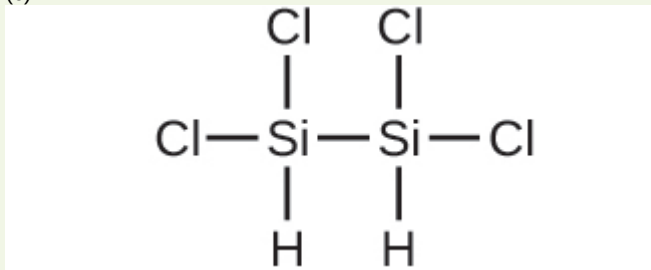
(a)



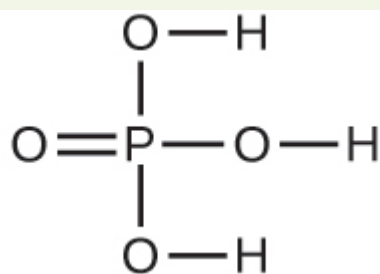
(b)



(c)



(d)

**Solution**

- (a) 56.107 amu;
- (b) 54.091 amu;
- (c) 199.9976 amu;
- (d) 97.9950 amu

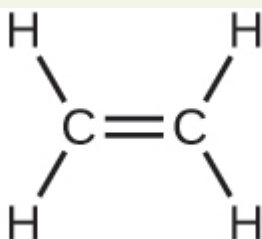
Exercise 1.23

Which molecule has a molecular mass of 28.05 amu?

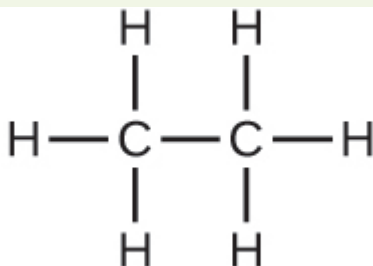
(a)



(b)



(c)

**Exercise 1.24**

Write a sentence that describes how to determine the number of moles of a compound in a known mass of the compound if we know its molecular formula.

Solution

Use the molecular formula to find the molar mass; to obtain the number of moles, divide the mass of compound by the molar mass of the compound expressed in grams.

Exercise 1.25

Compare 1 mole of H₂, 1 mole of O₂, and 1 mole of F₂.

- (a) Which has the largest number of molecules? Explain why.
- (b) Which has the greatest mass? Explain why.

Exercise 1.26

Which contains the greatest mass of oxygen: 0.75 mol of ethanol ($\text{C}_2\text{H}_5\text{OH}$), 0.60 mol of formic acid (HCO_2H), or 1.0 mol of water (H_2O)? Explain why.

Solution

Formic acid. Its formula has twice as many oxygen atoms as the other two compounds (one each). Therefore, 0.60 mol of formic acid would be equivalent to 1.20 mol of a compound containing a single oxygen atom.

Exercise 1.27

Which contains the greatest number of moles of oxygen atoms: 1 mol of ethanol ($\text{C}_2\text{H}_5\text{OH}$), 1 mol of formic acid (HCO_2H), or 1 mol of water (H_2O)? Explain why.

Exercise 1.28

How are the molecular mass and the molar mass of a compound similar and how are they different?

Solution

The two masses have the same numerical value, but the units are different: The molecular mass is the mass of 1 molecule while the molar mass is the mass of 6.022×10^{23} molecules.

Exercise 1.29

Calculate the molar mass of each of the following compounds:

- (a) hydrogen fluoride, HF
- (b) ammonia, NH_3
- (c) nitric acid, HNO_3
- (d) silver sulfate, Ag_2SO_4
- (e) boric acid, $\text{B}(\text{OH})_3$

Exercise 1.30

Calculate the molar mass of each of the following:

- (a) S_8
- (b) C_5H_{12}
- (c) $\text{Sc}_2(\text{SO}_4)_3$
- (d) CH_3COCH_3 (acetone)
- (e) $\text{C}_6\text{H}_{12}\text{O}_6$ (glucose)

Solution

(a) 256.528 g/mol; (b) 72.150 g mol⁻¹; (c) 378.103 g mol⁻¹; (d) 58.080 g mol⁻¹; (e) 180.158 g mol⁻¹

Exercise 1.31

Calculate the empirical or molecular formula mass and the molar mass of each of the following minerals:

- (a) limestone, CaCO_3
- (b) halite, NaCl
- (c) beryl, $\text{Be}_3\text{Al}_2\text{Si}_6\text{O}_{18}$
- (d) malachite, $\text{Cu}_2(\text{OH})_2\text{CO}_3$
- (e) turquoise, $\text{CuAl}_6(\text{PO}_4)_4(\text{OH})_8(\text{H}_2\text{O})_4$

Exercise 1.32

Calculate the molar mass of each of the following:

- (a) the anesthetic halothane, $\text{C}_2\text{HBrClF}_3$
- (b) the herbicide paraquat, $\text{C}_{12}\text{H}_{14}\text{N}_2\text{Cl}_2$
- (c) caffeine, $\text{C}_8\text{H}_{10}\text{N}_4\text{O}_2$
- (d) urea, $\text{CO}(\text{NH}_2)_2$
- (e) a typical soap, $\text{C}_{17}\text{H}_{35}\text{CO}_2\text{Na}$

Solution

(a) $197.382 \text{ g mol}^{-1}$; (b) $257.163 \text{ g mol}^{-1}$; (c) $194.193 \text{ g mol}^{-1}$; (d) $60.056 \text{ g mol}^{-1}$; (e) $306.464 \text{ g mol}^{-1}$

Exercise 1.33

Determine the number of moles of compound and the number of moles of each type of atom in each of the following:

- (a) 25.0 g of propylene, C_3H_6
- (b) $3.06 \times 10^{-3} \text{ g}$ of the amino acid glycine, $\text{C}_2\text{H}_5\text{NO}_2$
- (c) 25 lb of the herbicide Treflan, $\text{C}_{13}\text{H}_{16}\text{N}_2\text{O}_4\text{F}$ (1 lb = 454 g)
- (d) 0.125 kg of the insecticide Paris Green, $\text{Cu}_4(\text{AsO}_3)_2(\text{CH}_3\text{CO}_2)_2$
- (e) 325 mg of aspirin, $\text{C}_6\text{H}_4(\text{CO}_2\text{H})(\text{CO}_2\text{CH}_3)$

Exercise 1.34

Determine the mass of each of the following:

- (a) 0.0146 mol KOH
- (b) 10.2 mol ethane, C_2H_6
- (c) $1.6 \times 10^{-3} \text{ mol Na}_2\text{SO}_4$
- (d) $6.854 \times 10^3 \text{ mol}$ glucose, $\text{C}_6\text{H}_{12}\text{O}_6$
- (e) 2.86 mol $\text{Co}(\text{NH}_3)_6\text{Cl}_3$

Solution

- (a) 0.819 g;
- (b) 307 g;
- (c) 0.23 g;
- (d) $1.235 \times 10^6 \text{ g}$ (1235 kg);
- (e) 765 g

Exercise 1.35

Determine the number of moles of the compound and determine the number of moles of each type of atom in each of the following:

- (a) 2.12 g of potassium bromide, KBr
- (b) 0.1488 g of phosphoric acid, H_3PO_4
- (c) 23 kg of calcium carbonate, CaCO_3
- (d) 78.452 g of aluminum sulfate, $\text{Al}_2(\text{SO}_4)_3$
- (e) 0.1250 mg of caffeine, $\text{C}_8\text{H}_{10}\text{N}_4\text{O}_2$

Exercise 1.36

Determine the mass of each of the following:

- (a) 2.345 mol LiCl
- (b) 0.0872 mol acetylene, C_2H_2
- (c) $3.3 \times 10^{-2} \text{ mol Na}_2\text{CO}_3$

(d) 1.23×10^3 mol fructose, $C_6H_{12}O_6$

(e) 0.5758 mol $FeSO_4(H_2O)_7$

Solution

(a) 99.41 g;

(b) 2.27 g;

(c) 3.5 g;

(d) 222 kg;

(e) 160.1 g

Exercise 1.37

The approximate minimum daily dietary requirement of the amino acid leucine, $C_6H_{13}NO_2$, is 1.1 g. What is this requirement in moles?

Exercise 1.38

Determine the mass in grams of each of the following:

(a) 0.600 mol of oxygen atoms

(b) 0.600 mol of oxygen molecules, O_2

(c) 0.600 mol of ozone molecules, O_3

Solution

(a) 9.60 g; (b) 19.2 g; (c) 28.8 g

Exercise 1.39

A 55-kg woman has 7.5×10^{-3} mol of hemoglobin (molar mass = 64,456 g/mol) in her blood. How many hemoglobin molecules is this? What is this quantity in grams?

Exercise 1.40

Determine the number of atoms and the mass of zirconium, silicon, and oxygen found in 0.3384 mol of zircon, $ZrSiO_4$, a semiprecious stone.

Solution

zirconium: 2.038×10^{23} atoms; 30.87 g; silicon: 2.038×10^{23} atoms; 9.504 g; oxygen: 8.151×10^{23} atoms; 21.66 g

Exercise 1.41

Determine which of the following contains the greatest mass of hydrogen: 1 mol of CH_4 , 0.6 mol of C_6H_6 , or 0.4 mol of C_3H_8 .

Exercise 1.42

Determine which of the following contains the greatest mass of aluminum: 122 g of $AlPO_4$, 266 g of Al_2Cl_6 , or 225 g of Al_2S_3 .

Solution

$AlPO_4$: 1.000 mol, or 26.98 g Al

Al_2Cl_6 : 1.994 mol, or 53.74 g Al

Al_2S_3 : 3.00 mol, or 80.94 g Al

The Al_2S_3 sample thus contains the greatest mass of Al.

Exercise 1.43

Diamond is one form of elemental carbon. An engagement ring contains a diamond weighing 1.25 carats (1 carat = 200 mg). How many atoms are present in the diamond?

Exercise 1.44

The Cullinan diamond was the largest natural diamond ever found (January 25, 1905). It weighed 3104 carats (1 carat = 200 mg). How many carbon atoms were present in the stone?

Solution

$$3.113 \times 10^{25} \text{ C atoms}$$

Exercise 1.45

One 55-gram serving of a particular cereal supplies 270 mg of sodium, 11% of the recommended daily allowance. How many moles and atoms of sodium are in the recommended daily allowance?

Exercise 1.46

A certain nut crunch cereal contains 11.0 grams of sugar (sucrose, $\text{C}_{12}\text{H}_{22}\text{O}_{11}$) per serving size of 60.0 grams. How many servings of this cereal must be eaten to consume 0.0278 moles of sugar?

Solution

0.865 servings, or about 1 serving.

Exercise 1.47

A tube of toothpaste contains 0.76 g of sodium monofluorophosphate ($\text{Na}_2\text{PO}_3\text{F}$) in 100 mL.

- (a) What mass of fluorine atoms in mg was present?
- (b) How many fluorine atoms were present?

Exercise 1.48

Which of the following represents the least number of molecules?

- (a) 20.0 g of H_2O (18.02 g/mol)
- (b) 77.0 g of CH_4 (16.06 g/mol)
- (c) 68.0 g of CaH_2 (42.09 g/mol)
- (d) 100.0 g of N_2O (44.02 g/mol)
- (e) 84.0 g of HF (20.01 g/mol)

Solution

20.0 g H_2O represents the least number of molecules since it has the least number of moles.

1.5 The Molecular Picture of Gasses

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The section on ideal gasses vs. real gasses is also available as a video [here \(https://www.youtube.com/watch?v=3vGfWoPlstk\)](https://www.youtube.com/watch?v=3vGfWoPlstk).



Figure 1.17 The air inside this hot air balloon flying over Putrajaya, Malaysia, is hotter than the ambient air. As a result, the balloon experiences a buoyant force pushing it upward. (credit: Kevin Poh, Flickr)

In this section, we continue to explore the thermal behavior of gases. In particular, we examine the characteristics of atoms and molecules that compose gases. (Most gases, for example nitrogen, N_2 , and oxygen, O_2 , are composed of two or more atoms. We will primarily use the term “molecule” in discussing a gas because the term can also be applied to monatomic gases, such as helium.)

Gases are easily compressed. We can see evidence of this in [m42215 \(https://legacy.cnx.org/content/m42215/latest/#import-auto-id1814176\)](https://legacy.cnx.org/content/m42215/latest/#import-auto-id1814176), where you will note that gases have the *largest* coefficients of volume expansion. The large coefficients mean that gases expand and contract very rapidly with temperature changes. In addition, you will note that most gases expand at the *same* rate, or have the same β . This raises the question as to why gases should all act in nearly the same way, when liquids and solids have widely varying expansion rates.

The answer lies in the large separation of atoms and molecules in gases, compared to their sizes, as illustrated in **Figure 1.18**. Because atoms and molecules have large separations, forces between them can be ignored, except when they collide with each other during collisions. The motion of atoms and molecules (at temperatures well above the boiling temperature) is fast, such that the gas occupies all of the accessible volume and the expansion of gases is rapid. In contrast, in liquids and solids, atoms and molecules are closer together and are quite sensitive to the forces between them.

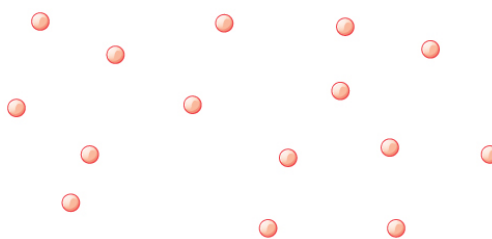


Figure 1.18 Atoms and molecules in a gas are typically widely separated, as shown. Because the forces between them are quite weak at these distances, the properties of a gas depend more on the number of atoms per unit volume and on temperature than on the type of atom.

To get some idea of how pressure, temperature, and volume of a gas are related to one another, consider what happens when you pump air into an initially deflated tire. The tire’s volume first increases in direct proportion to the amount of air injected, without much increase in the tire pressure. Once the tire has expanded to nearly its full size, the walls limit volume expansion. If we continue to pump air into it, the pressure increases. The pressure will further increase when the car is driven and the tires move. Most manufacturers specify optimal tire pressure for cold tires. (See **Figure 1.19**.)

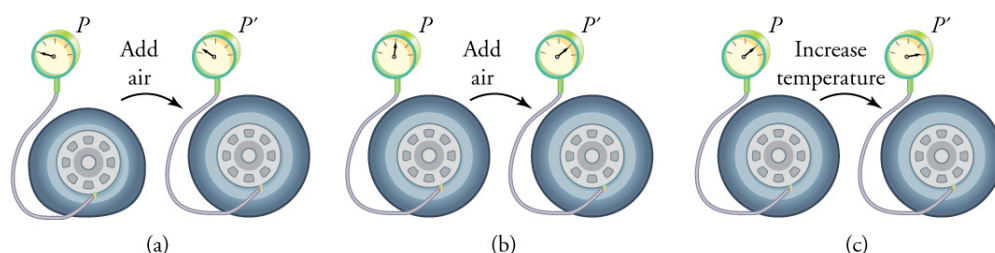


Figure 1.19 (a) When air is pumped into a deflated tire, its volume first increases without much increase in pressure. (b) When the tire is filled to a certain point, the tire walls resist further expansion and the pressure increases with more air. (c) Once the tire is inflated, its pressure increases with temperature.

At room temperatures, collisions between atoms and molecules can be ignored. In this case, the gas is called an ideal gas, in which case the relationship between the pressure, volume, and temperature is given by the equation of state called the ideal gas law.

Ideal Gas Law

The **ideal gas law** states that

$$PV = NkT, \quad (1.14)$$

where P is the absolute pressure of a gas, V is the volume it occupies, N is the number of atoms and molecules in the gas, and T is its absolute temperature. The constant k is called the **Boltzmann constant** in honor of Austrian physicist Ludwig Boltzmann (1844–1906) and has the value

$$k = 1.38 \times 10^{-23} \text{ J/K}. \quad (1.15)$$

The ideal gas law can be derived from basic principles, but was originally deduced from experimental measurements of Charles' law (that volume occupied by a gas is proportional to temperature at a fixed pressure) and from Boyle's law (that for a fixed temperature, the product PV is a constant). In the ideal gas model, the volume occupied by its atoms and molecules is a negligible fraction of V . The ideal gas law describes the behavior of real gases under most conditions. (Note, for example, that N is the total number of atoms and molecules, independent of the type of gas.)

Let us see how the ideal gas law is consistent with the behavior of filling the tire when it is pumped slowly and the temperature is constant. At first, the pressure P is essentially equal to atmospheric pressure, and the volume V increases in direct proportion to the number of atoms and molecules N put into the tire. Once the volume of the tire is constant, the equation $PV = NkT$ predicts that the pressure should increase in proportion to the number N of atoms and molecules.

Example 1.12 Calculating Pressure Changes Due to Temperature Changes: Tire Pressure

Suppose your bicycle tire is fully inflated, with an absolute pressure of $7.00 \times 10^5 \text{ Pa}$ (a gauge pressure of just under 90.0 lb/in^2) at a temperature of 18.0°C . What is the pressure after its temperature has risen to 35.0°C ? Assume that there are no appreciable leaks or changes in volume.

Strategy

The pressure in the tire is changing only because of changes in temperature. First we need to identify what we know and what we want to know, and then identify an equation to solve for the unknown.

We know the initial pressure $P_0 = 7.00 \times 10^5 \text{ Pa}$, the initial temperature $T_0 = 18.0^\circ\text{C}$, and the final temperature $T_f = 35.0^\circ\text{C}$. We must find the final pressure P_f . How can we use the equation $PV = NkT$? At first, it may seem that not enough information is given, because the volume V and number of atoms N are not specified. What we can do is use the equation twice: $P_0 V_0 = NkT_0$ and $P_f V_f = NkT_f$. If we divide $P_f V_f$ by $P_0 V_0$ we can come up with an equation that allows us to solve for P_f .

$$\frac{P_f V_f}{P_0 V_0} = \frac{N_f k T_f}{N_0 k T_0} \quad (1.16)$$

Since the volume is constant, V_f and V_0 are the same and they cancel out. The same is true for N_f and N_0 , and k , which is a constant. Therefore,

$$\frac{P_f}{P_0} = \frac{T_f}{T_0}. \quad (1.17)$$

We can then rearrange this to solve for P_f :

$$P_f = P_0 \frac{T_f}{T_0}, \quad (1.18)$$

where the temperature must be in units of kelvins, because T_0 and T_f are absolute temperatures.

Solution

1. Convert temperatures from Celsius to Kelvin.

$$T_0 = (18.0 + 273)\text{K} = 291 \text{ K} \quad (1.19)$$

$$T_f = (35.0 + 273)\text{K} = 308 \text{ K}$$

2. Substitute the known values into the equation.

$$P_f = P_0 \frac{T_f}{T_0} = 7.00 \times 10^5 \text{ Pa} \left(\frac{308 \text{ K}}{291 \text{ K}} \right) = 7.41 \times 10^5 \text{ Pa} \quad (1.20)$$

Discussion

The final temperature is about 6% greater than the original temperature, so the final pressure is about 6% greater as well. Note that *absolute* pressure and *absolute* temperature must be used in the ideal gas law.

Making Connections: Take-Home Experiment—Refrigerating a Balloon

Inflate a balloon at room temperature. Leave the inflated balloon in the refrigerator overnight. What happens to the balloon, and why?

Example 1.13 Calculating the Number of Molecules in a Cubic Meter of Gas

How many molecules are in a typical object, such as gas in a tire or water in a drink? We can use the ideal gas law to give us an idea of how large N typically is.

Calculate the number of molecules in a cubic meter of gas at standard temperature and pressure (STP), which is defined to be 0°C and atmospheric pressure.

Strategy

Because pressure, volume, and temperature are all specified, we can use the ideal gas law $PV = NkT$, to find N .

Solution

1. Identify the knowns.

$$T = 0^\circ\text{C} = 273 \text{ K} \quad (1.21)$$

$$P = 1.01 \times 10^5 \text{ Pa}$$

$$V = 1.00 \text{ m}^3$$

$$k = 1.38 \times 10^{-23} \text{ J/K}$$

2. Identify the unknown: number of molecules, N .

3. Rearrange the ideal gas law to solve for N .

$$PV = NkT \quad (1.22)$$

$$N = \frac{PV}{kT}$$

4. Substitute the known values into the equation and solve for N .

$$N = \frac{PV}{kT} = \frac{(1.01 \times 10^5 \text{ Pa})(1.00 \text{ m}^3)}{(1.38 \times 10^{-23} \text{ J/K})(273 \text{ K})} = 2.68 \times 10^{25} \text{ molecules} \quad (1.23)$$

Discussion

This number is undeniably large, considering that a gas is mostly empty space. N is huge, even in small volumes. For example, 1 cm^3 of a gas at STP has 2.68×10^{19} molecules in it. Once again, note that N is the same for all types or mixtures of gases.

The Ideal Gas Law Restated Using Moles

A very common expression of the ideal gas law uses the number of moles, n , rather than the number of atoms and molecules, N . We start from the ideal gas law,

$$PV = NkT, \quad (1.24)$$

and multiply and divide the equation by Avogadro's number N_A . This gives

$$PV = \frac{N}{N_A} N_A kT. \quad (1.25)$$

Note that $n = N/N_A$ is the number of moles. We define the universal gas constant $R = N_A k$, and obtain the ideal gas law in terms of moles.

Ideal Gas Law (in terms of moles)

The ideal gas law (in terms of moles) is

$$PV = nRT. \quad (1.26)$$

The numerical value of R in SI units is

$$R = N_A k = (6.02 \times 10^{23} \text{ mol}^{-1})(1.38 \times 10^{-23} \text{ J/K}) = 8.31 \text{ J/mol} \cdot \text{K}. \quad (1.27)$$

In other units,

$$\begin{aligned} R &= 1.99 \text{ cal/mol} \cdot \text{K} \\ R &= 0.0821 \text{ L} \cdot \text{atm/mol} \cdot \text{K}. \end{aligned} \quad (1.28)$$

You can use whichever value of R is most convenient for a particular problem.

Example 1.14 Calculating Number of Moles: Gas in a Bike Tire

How many moles of gas are in a bike tire with a volume of $2.00 \times 10^{-3} \text{ m}^3$ (2.00 L), a pressure of $7.00 \times 10^5 \text{ Pa}$ (a gauge pressure of just under 90.0 lb/in^2), and at a temperature of 18.0°C ?

Strategy

Identify the knowns and unknowns, and choose an equation to solve for the unknown. In this case, we solve the ideal gas law, $PV = nRT$, for the number of moles n .

Solution

1. Identify the knowns.

$$\begin{aligned} P &= 7.00 \times 10^5 \text{ Pa} \\ V &= 2.00 \times 10^{-3} \text{ m}^3 \\ T &= 18.0^\circ\text{C} = 291 \text{ K} \\ R &= 8.31 \text{ J/mol} \cdot \text{K} \end{aligned} \quad (1.29)$$

2. Rearrange the equation to solve for n and substitute known values.

$$\begin{aligned} n &= \frac{PV}{RT} = \frac{(7.00 \times 10^5 \text{ Pa})(2.00 \times 10^{-3} \text{ m}^3)}{(8.31 \text{ J/mol} \cdot \text{K})(291 \text{ K})} \\ &= 0.579 \text{ mol} \end{aligned} \quad (1.30)$$

Discussion

The most convenient choice for R in this case is $8.31 \text{ J/mol} \cdot \text{K}$, because our known quantities are in SI units. The pressure and temperature are obtained from the initial conditions in **Example 1.12**, but we would get the same answer if we

used the final values.

The ideal gas law can be considered to be another manifestation of the law of conservation of energy (see **Conservation of Energy** (<https://legacy.cnx.org/content/m42151/latest/>)). Work done on a gas results in an increase in its energy, increasing pressure and/or temperature, or decreasing volume. This increased energy can also be viewed as increased internal kinetic energy, given the gas's atoms and molecules.

Ideal Gasses vs Real Gasses

To understand the ideal gas approximation, it's probably best to compare it to a real gas, such as the oxygen or nitrogen in the room around you. In a real gas, the gas is made up of particles with some small, but not exactly zero volume. In the ideal gas approximation, we assume that the molecules have exactly zero size. Similarly, in a real gas the molecules attract each other ever so slightly through van der Waals forces and the like. An ideal gas molecule, however, does not attract another ideal gas molecule, they don't interact with each other at all. Also in our analysis of the ideal gas, we will typically ignore gravity. We'll just turn it off and pretend it doesn't have an effect.

These assumptions may seem to have no connection to reality. I mean, a gas of particles of zero size that don't interact with each other and have no impact by gravity? Those seem like fairly, you know, unphysical conditions. However, if you think about an oxygen molecule, its volume is 10^{-30} m^3 , which is very small compared to the room you're in, so the approximation that a gas has zero size for its particles is not actually that bad. Similarly, the force of the hydrogen bond is something on the neighborhood of 10^{-10} N , which again, is a very small force, so the idea that ideal gas molecules don't attract each other at all is not a terrible approximation. In fact, carbon dioxide gas behaves like an ideal gas to 0.5% accuracy at everyday pressures and temperatures. So, we can use the ideal gas to get a very good idea of how real gases are behaving under a variety of typical conditions.

Let's explore the consequences of this ideal gas approximation. So first we've turned off gravity, what effect does that have? Well, without gravity the particles have no gravitational potential energy, only kinetic energies are possible. There's no g for the mgh . Furthermore, without gravity, particles travel in a straight line until they run into something. They do not follow the parabolic arcs that we've been studying in our third unit. Now let's explore the consequences of the fact that ideal gas particles have zero size and zero intermolecular force. Well, two particles of zero size cannot collide. Think about that one for a minute, but you'll see it's true. Since they can't collide with each other, the particles only hit the walls. The combination of these two approximations means that ideal gas particles travel in straight lines until they hit a wall.

Section Summary

- The ideal gas law relates the pressure and volume of a gas to the number of gas molecules and the temperature of the gas.
- The ideal gas law can be written in terms of the number of molecules of gas:

$$PV = NkT, \quad (1.31)$$

where P is pressure, V is volume, T is temperature, N is number of molecules, and k is the Boltzmann constant

$$k = 1.38 \times 10^{-23} \text{ J/K}. \quad (1.32)$$

- A mole is the number of atoms in a 12-g sample of carbon-12.
- The number of molecules in a mole is called Avogadro's number N_A ,

$$N_A = 6.02 \times 10^{23} \text{ mol}^{-1}. \quad (1.33)$$

- A mole of any substance has a mass in grams equal to its molecular weight, which can be determined from the periodic table of elements.
- The ideal gas law can also be written and solved in terms of the number of moles of gas:

$$PV = nRT, \quad (1.34)$$

where n is number of moles and R is the universal gas constant,

$$R = 8.31 \text{ J/mol} \cdot \text{K}. \quad (1.35)$$

- The ideal gas law is generally valid at temperatures well above the boiling temperature.

Conceptual Questions

Exercise 1.49

Find out the human population of Earth. Is there a mole of people inhabiting Earth? If the average mass of a person is 60 kg, calculate the mass of a mole of people. How does the mass of a mole of people compare with the mass of Earth?

Exercise 1.50

Under what circumstances would you expect a gas to behave significantly differently than predicted by the ideal gas law?

Exercise 1.51

A constant-volume gas thermometer contains a fixed amount of gas. What property of the gas is measured to indicate its temperature?

Problems & Exercises**Exercise 1.52**

The gauge pressure in your car tires is $2.50 \times 10^5 \text{ N/m}^2$ at a temperature of 35.0°C when you drive it onto a ferry boat to Alaska. What is their gauge pressure later, when their temperature has dropped to -40.0°C ?

Solution
1.62 atm

Exercise 1.53

Convert an absolute pressure of $7.00 \times 10^5 \text{ N/m}^2$ to gauge pressure in lb/in^2 . (This value was stated to be just less than 90.0 lb/in^2 in **Example 1.14**. Is it?)

Exercise 1.54

Suppose a gas-filled incandescent light bulb is manufactured so that the gas inside the bulb is at atmospheric pressure when the bulb has a temperature of 20.0°C . (a) Find the gauge pressure inside such a bulb when it is hot, assuming its average temperature is 60.0°C (an approximation) and neglecting any change in volume due to thermal expansion or gas leaks. (b) The actual final pressure for the light bulb will be less than calculated in part (a) because the glass bulb will expand. What will the actual final pressure be, taking this into account? Is this a negligible difference?

Solution
(a) 0.136 atm
(b) 0.135 atm. The difference between this value and the value from part (a) is negligible.

Exercise 1.55

Large helium-filled balloons are used to lift scientific equipment to high altitudes. (a) What is the pressure inside such a balloon if it starts out at sea level with a temperature of 10.0°C and rises to an altitude where its volume is twenty times the original volume and its temperature is -50.0°C ? (b) What is the gauge pressure? (Assume atmospheric pressure is constant.)

Exercise 1.56

Confirm that the units of nRT are those of energy for each value of R : (a) $8.31 \text{ J/mol} \cdot \text{K}$, (b) $1.99 \text{ cal/mol} \cdot \text{K}$, and (c) $0.0821 \text{ L} \cdot \text{atm/mol} \cdot \text{K}$.

Solution

- (a) $nRT = (\text{mol})(\text{J/mol} \cdot \text{K})(\text{K}) = \text{J}$
 (b) $nRT = (\text{mol})(\text{cal/mol} \cdot \text{K})(\text{K}) = \text{cal}$
 $nRT = (\text{mol})(\text{L} \cdot \text{atm/mol} \cdot \text{K})(\text{K})$
 (c) $= \text{L} \cdot \text{atm} = (\text{m}^3)(\text{N/m}^2)$
 $= \text{N} \cdot \text{m} = \text{J}$

Exercise 1.57

In the text, it was shown that $N/V = 2.68 \times 10^{25} \text{ m}^{-3}$ for gas at STP. (a) Show that this quantity is equivalent to $N/V = 2.68 \times 10^{19} \text{ cm}^{-3}$, as stated. (b) About how many atoms are there in one μm^3 (a cubic micrometer) at STP? (c) What does your answer to part (b) imply about the separation of atoms and molecules?

Exercise 1.58

Calculate the number of moles in the 2.00-L volume of air in the lungs of the average person. Note that the air is at 37.0°C (body temperature).

Solution

$$7.86 \times 10^{-2} \text{ mol}$$

Exercise 1.59

An airplane passenger has 100 cm^3 of air in his stomach just before the plane takes off from a sea-level airport. What volume will the air have at cruising altitude if cabin pressure drops to $7.50 \times 10^4 \text{ N/m}^2$?

Exercise 1.60

(a) What is the volume (in km^3) of Avogadro's number of sand grains if each grain is a cube and has sides that are 1.0 mm long? (b) How many kilometers of beaches in length would this cover if the beach averages 100 m in width and 10.0 m in depth? Neglect air spaces between grains.

Solution

(a) $6.02 \times 10^5 \text{ km}^3$

(b) $6.02 \times 10^8 \text{ km}$

Exercise 1.61

An expensive vacuum system can achieve a pressure as low as $1.00 \times 10^{-7} \text{ N/m}^2$ at 20°C . How many atoms are there in a cubic centimeter at this pressure and temperature?

Exercise 1.62

The number density of gas atoms at a certain location in the space above our planet is about $1.00 \times 10^{11} \text{ m}^{-3}$, and the pressure is $2.75 \times 10^{-10} \text{ N/m}^2$ in this space. What is the temperature there?

Solution

$$-73.9^\circ\text{C}$$

Exercise 1.63

A bicycle tire has a pressure of $7.00 \times 10^5 \text{ N/m}^2$ at a temperature of 18.0°C and contains 2.00 L of gas. What will its pressure be if you let out an amount of air that has a volume of 100 cm^3 at atmospheric pressure? Assume tire temperature and volume remain constant.

Exercise 1.64

A high-pressure gas cylinder contains 50.0 L of toxic gas at a pressure of $1.40 \times 10^7 \text{ N/m}^2$ and a temperature of 25.0°C .

Its valve leaks after the cylinder is dropped. The cylinder is cooled to dry ice temperature (-78.5°C) to reduce the leak rate and pressure so that it can be safely repaired. (a) What is the final pressure in the tank, assuming a negligible amount of gas leaks while being cooled and that there is no phase change? (b) What is the final pressure if one-tenth of the gas escapes? (c) To what temperature must the tank be cooled to reduce the pressure to 1.00 atm (assuming the gas does not change phase and that there is no leakage during cooling)? (d) Does cooling the tank appear to be a practical solution?

Solution

- (a) $9.14 \times 10^6 \text{ N/m}^2$
- (b) $8.23 \times 10^6 \text{ N/m}^2$
- (c) 2.16 K
- (d) No. The final temperature needed is much too low to be easily achieved for a large object.

Exercise 1.65

Find the number of moles in 2.00 L of gas at 35.0°C and under $7.41 \times 10^7 \text{ N/m}^2$ of pressure.

Exercise 1.66

Calculate the depth to which Avogadro's number of table tennis balls would cover Earth. Each ball has a diameter of 3.75 cm. Assume the space between balls adds an extra 25.0% to their volume and assume they are not crushed by their own weight.

Solution

41 km

Exercise 1.67

(a) What is the gauge pressure in a 25.0°C car tire containing 3.60 mol of gas in a 30.0 L volume? (b) What will its gauge pressure be if you add 1.00 L of gas originally at atmospheric pressure and 25.0°C ? Assume the temperature returns to 25.0°C and the volume remains constant.

Exercise 1.68

(a) In the deep space between galaxies, the density of atoms is as low as 10^6 atoms/m^3 , and the temperature is a frigid 2.7 K. What is the pressure? (b) What volume (in m^3) is occupied by 1 mol of gas? (c) If this volume is a cube, what is the length of its sides in kilometers?

Solution

- (a) $3.7 \times 10^{-17} \text{ Pa}$
- (b) $6.0 \times 10^{17} \text{ m}^3$
- (c) $8.4 \times 10^2 \text{ km}$

1.6 The Molecular Picture of Solids

What is a Solid?

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Instructor's Notes

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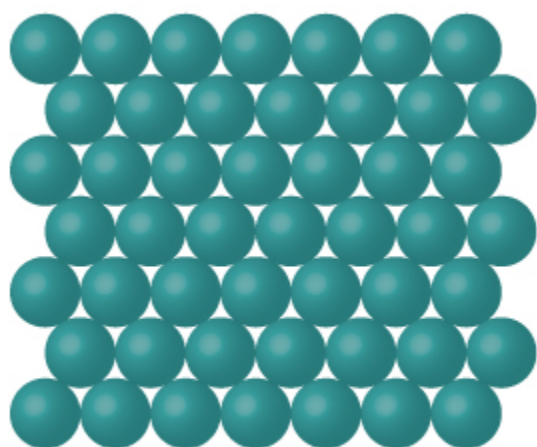
Much of the matter we typically deal with in our everyday lives is a solid. The idea of "solid" is something that retains its shape. This is an idealization -- it really means "it doesn't change its shape while I'm watching it." As we know, a solid can change its shape if you push on it hard enough (by deforming or breaking). Some things that appear solid will change their shape slowly under the gentle pull of gravity. Some glass in very old windows shows a small amount of sagging. And what is it with butter, anyway?

A *uniform solid* is one where every part of it is the same as every other part. This is a model statement, since all matter is composed of atoms and at the atomic level no matter is uniform since there are spaces between the atoms. When we say something is a "uniform solid" we mean that we are going to be examining it on a scale large enough that we can ignore atomic discontinuities.

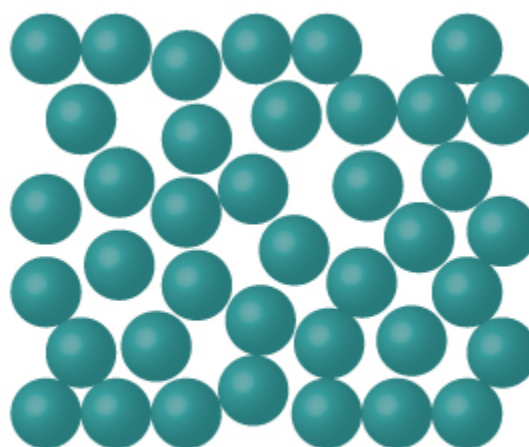
Some uniform solids consist of atoms or ions in a regular pattern, such as diamonds composed only of carbon atoms in a regular lattice, or a salt crystal composed of sodium and chloride ions in a regular array. However, these are relatively uncommon and so for the most part we will consider solids to be made up of molecules. Solids can be formed from a single compound where all the molecules are the same, or multiple compounds, with a mixture of molecules. A multi-component solid is sometimes harder to characterize as it can be a non-uniform mixture or conglomerate.

Biological solids can be quite complex. They are often not uniform, but instead are composites, made of particles embedded in layers composed of different materials. Your skin is a good example. It is composed of a thick layer of connective tissue (dermis) underlain by a membrane and overlaid by a surface (epidermal) layer. The dermis is actually a three dimensional network of collagen fibers that are embedded in a protein-polysaccharide matrix. The stretchiness of your skin is provided by additional elastin fibers which are distributed throughout the dermis.*

When most liquids are cooled, they eventually freeze and form **crystalline solids**, solids in which the atoms, ions, or molecules are arranged in a definite repeating pattern. It is also possible for a liquid to freeze before its molecules become arranged in an orderly pattern. The resulting materials are called **amorphous solids** or noncrystalline solids (or, sometimes, glasses). The particles of such solids lack an ordered internal structure and are randomly arranged (**Figure 1.20**).



Crystalline



Amorphous

Figure 1.20 The entities of a solid phase may be arranged in a regular, repeating pattern (crystalline solids) or randomly (amorphous).

Metals and ionic compounds typically form ordered, crystalline solids. Substances that consist of large molecules, or a mixture of molecules whose movements are more restricted, often form amorphous solids. For examples, candle waxes are amorphous solids composed of large hydrocarbon molecules. Some substances, such as boron oxide (shown in **Figure 1.21**), can form either crystalline or amorphous solids, depending on the conditions under which it is produced. Also, amorphous solids may undergo a transition to the crystalline state under appropriate conditions.

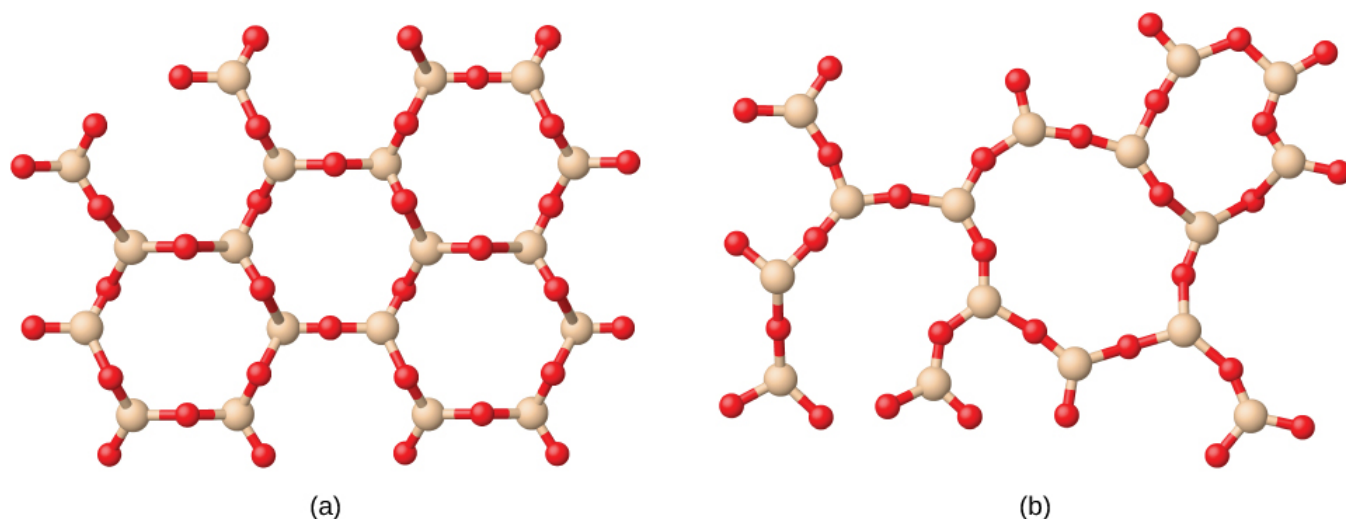


Figure 1.21 (a) Diboron trioxide, B_2O_3 , is normally found as a white, amorphous solid (a glass), which has a high degree of disorder in its structure. (b) By careful, extended heating, it can be converted into a crystalline form of B_2O_3 , which has a very ordered arrangement.

Crystalline solids are generally classified according to the nature of the forces that hold its particles together. These forces are primarily responsible for the physical properties exhibited by the bulk solids. The following sections provide descriptions of the major types of crystalline solids: ionic, metallic, covalent network, and molecular.

Ionic Solids

Ionic solids, such as sodium chloride and nickel oxide, are composed of positive and negative ions that are held together by electrostatic attractions, which can be quite strong (**Figure 1.22**). Many ionic crystals also have high melting points. This is due to the very strong attractions between the ions—in ionic compounds, the attractions between full charges are (much) larger than those between the partial charges in polar molecular compounds. This will be looked at in more detail in a later discussion of lattice energies. Although they are hard, they also tend to be brittle, and they shatter rather than bend. Ionic solids do not conduct electricity; however, they do conduct when molten or dissolved because their ions are free to move. Many simple compounds formed by the reaction of a metallic element with a nonmetallic element are ionic.

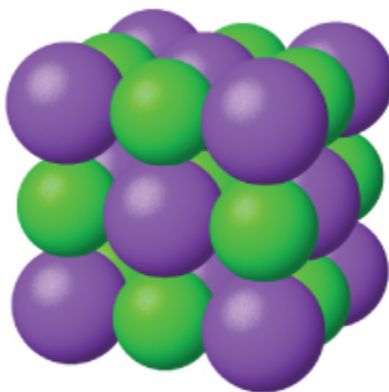


Figure 1.22 Sodium chloride is an ionic solid.

Metallic Solids

Metallic solids such as crystals of copper, aluminum, and iron are formed by metal atoms **Figure 1.23**. The structure of metallic crystals is often described as a uniform distribution of atomic nuclei within a “sea” of delocalized electrons. The atoms within such a metallic solid are held together by a unique force known as *metallic bonding* that gives rise to many useful and varied bulk properties. All exhibit high thermal and electrical conductivity, metallic luster, and malleability. Many are very hard and quite strong. Because of their malleability (the ability to deform under pressure or hammering), they do not shatter and, therefore, make useful construction materials. The melting points of the metals vary widely. Mercury is a liquid at room temperature, and the alkali metals melt below $200^\circ C$. Several post-transition metals also have low melting points, whereas the transition metals melt at temperatures above $1000^\circ C$. These differences reflect differences in strengths of metallic bonding among the metals.

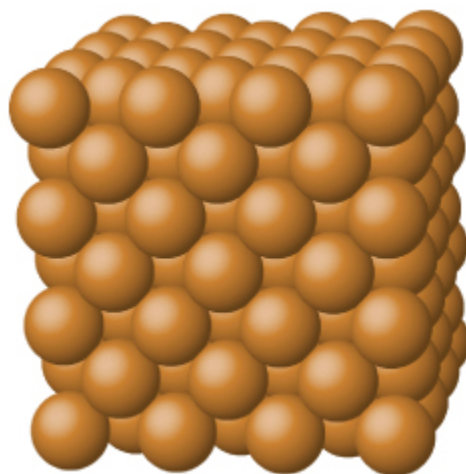
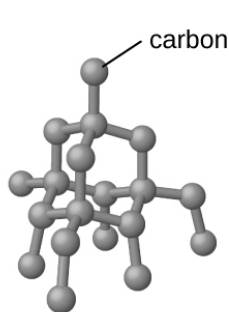
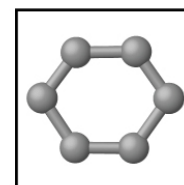
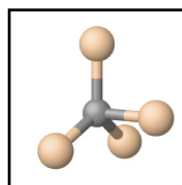
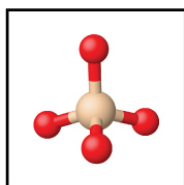
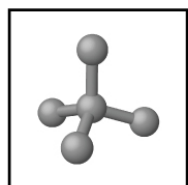


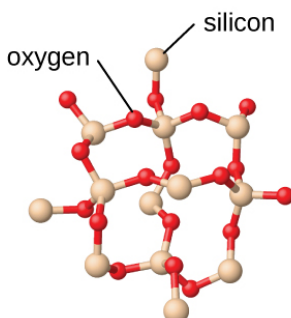
Figure 1.23 Copper is a metallic solid.

Covalent Network Solid

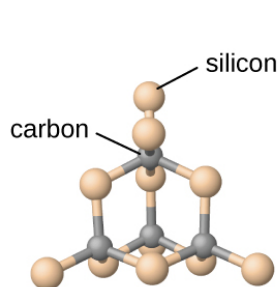
Covalent network solids include crystals of diamond, silicon, some other nonmetals, and some covalent compounds such as silicon dioxide (sand) and silicon carbide (carborundum, the abrasive on sandpaper). Many minerals have networks of covalent bonds. The atoms in these solids are held together by a network of covalent bonds, as shown in Figure 1.24. To break or to melt a covalent network solid, covalent bonds must be broken. Because covalent bonds are relatively strong, covalent network solids are typically characterized by hardness, strength, and high melting points. For example, diamond is one of the hardest substances known and melts above 3500 °C.



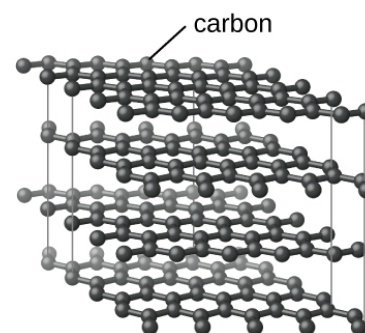
diamond



silicon dioxide



silicon carbide



graphite

Figure 1.24 A covalent crystal contains a three-dimensional network of covalent bonds, as illustrated by the structures of diamond, silicon dioxide, silicon carbide, and graphite. Graphite is an exceptional example, composed of planar sheets of covalent crystals that are held together in layers by noncovalent forces. Unlike typical covalent solids, graphite is very soft and electrically conductive.

Molecular Solid

Molecular solids, such as ice, sucrose (table sugar), and iodine, as shown in Figure 1.25, are composed of neutral molecules. The strengths of the attractive forces between the units present in different crystals vary widely, as indicated by the melting points of the crystals. Small symmetrical molecules (nonpolar molecules), such as H_2 , N_2 , O_2 , and F_2 , have weak attractive forces and form molecular solids with very low melting points (below $-200\text{ }^\circ\text{C}$). Substances consisting of larger, nonpolar molecules have larger attractive forces and melt at higher temperatures. Molecular solids composed of molecules with permanent dipole moments (polar molecules) melt at still higher temperatures. Examples include ice (melting point, $0\text{ }^\circ\text{C}$) and table sugar (melting point, $185\text{ }^\circ\text{C}$).

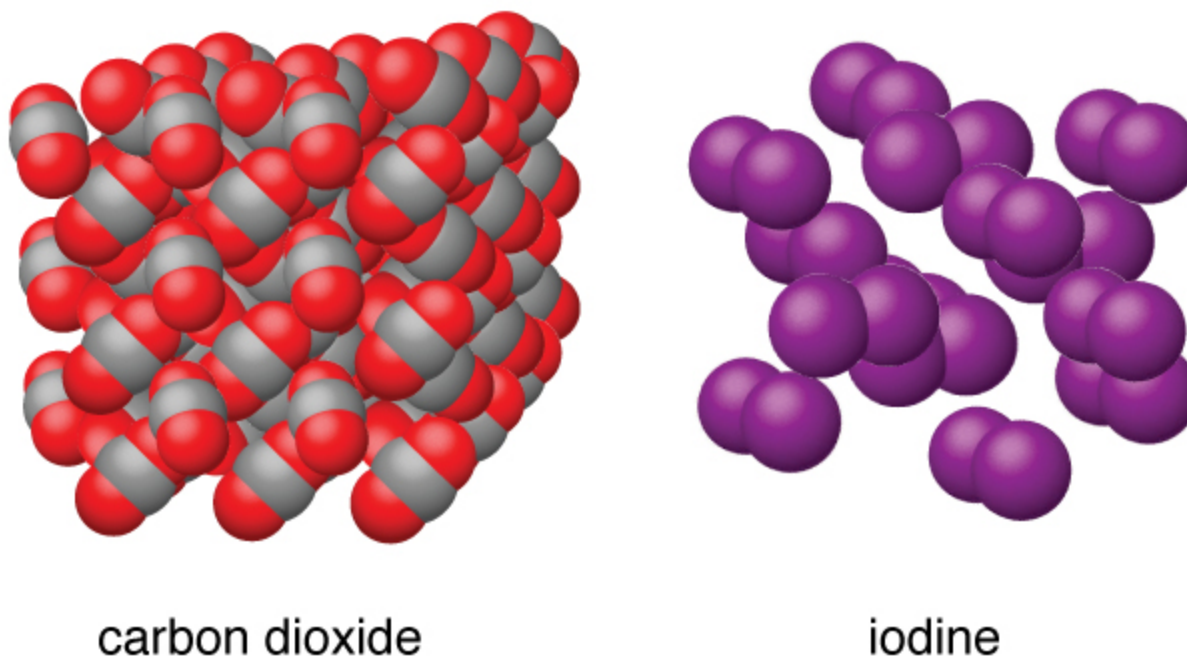


Figure 1.25 Carbon dioxide (CO_2) consists of small, nonpolar molecules and forms a molecular solid with a melting point of -78°C . Iodine (I_2) consists of larger, nonpolar molecules and forms a molecular solid that melts at 114°C .

Properties of Solids

A crystalline solid, like those listed in **Table 1.5**, has a precise melting temperature because each atom or molecule of the same type is held in place with the same forces or energy. Thus, the attractions between the units that make up the crystal all have the same strength and all require the same amount of energy to be broken. The gradual softening of an amorphous material differs dramatically from the distinct melting of a crystalline solid. This results from the structural nonequivalence of the molecules in the amorphous solid. Some forces are weaker than others, and when an amorphous material is heated, the weakest intermolecular attractions break first. As the temperature is increased further, the stronger attractions are broken. Thus amorphous materials soften over a range of temperatures.

Table 1.5

Types of Crystalline Solids and Their Properties				
Type of Solid	Type of Particles	Type of Attractions	Properties	Examples
ionic	ions	ionic bonds	hard, brittle, conducts electricity as a liquid but not as a solid, high to very high melting points	NaCl , Al_2O_3
metallic	atoms of electropositive elements	metallic bonds	shiny, malleable, ductile, conducts heat and electricity well, variable hardness and melting temperature	Cu , Fe , Ti , Pb , U
covalent network	atoms of electronegative elements	covalent bonds	very hard, not conductive, very high melting points	C (diamond), SiO_2 , SiC
molecular	molecules (or atoms)	IMFs	variable hardness, variable brittleness, not conductive, low melting points	H_2O , CO_2 , I_2 , $\text{C}_{12}\text{H}_{22}\text{O}_{11}$

Graphene: Material of the Future

Carbon is an essential element in our world. The unique properties of carbon atoms allow the existence of carbon-based life forms such as ourselves. Carbon forms a huge variety of substances that we use on a daily basis, including those shown in **Figure 1.26**. You may be familiar with diamond and graphite, the two most common *allotropes* of carbon. (Allotropes are different structural forms of the same element.) Diamond is one of the hardest-known substances, whereas graphite is soft enough to be used as pencil lead. These very different properties stem from the different arrangements of the carbon atoms in the different allotropes.

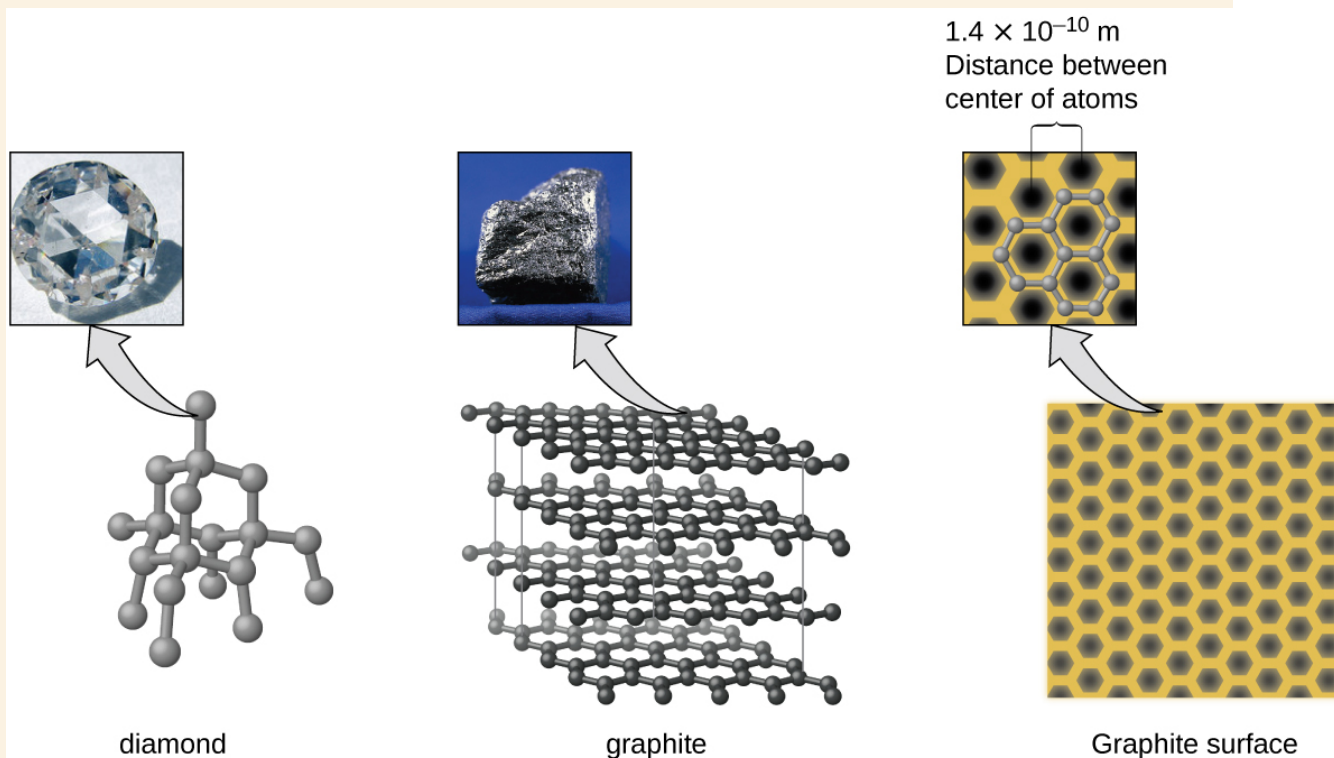


Figure 1.26 Diamond is extremely hard because of the strong bonding between carbon atoms in all directions. Graphite (in pencil lead) rubs off onto paper due to the weak attractions between the carbon layers. An image of a graphite surface shows the distance between the centers of adjacent carbon atoms. (credit left photo: modification of work by Steve Jurvetson; credit middle photo: modification of work by United States Geological Survey)

You may be less familiar with a recently discovered form of carbon: graphene. Graphene was first isolated in 2004 by using tape to peel off thinner and thinner layers from graphite. It is essentially a single sheet (one atom thick) of graphite.

Graphene, illustrated in **Figure 1.27**, is not only strong and lightweight, but it is also an excellent conductor of electricity and heat. These properties may prove very useful in a wide range of applications, such as vastly improved computer chips and circuits, better batteries and solar cells, and stronger and lighter structural materials. The 2010 Nobel Prize in Physics was awarded to Andre Geim and Konstantin Novoselov for their pioneering work with graphene.

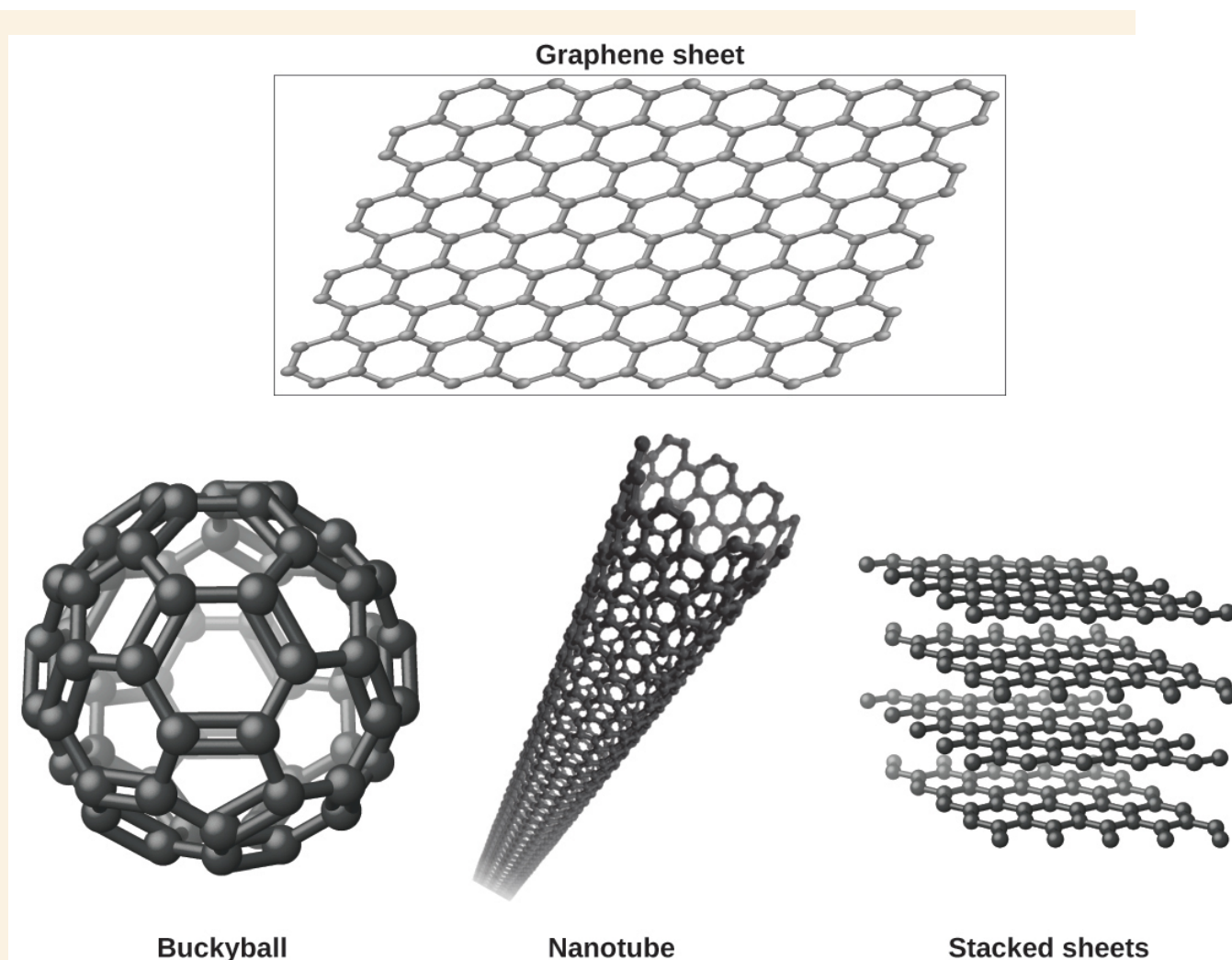


Figure 1.27 Graphene sheets can be formed into buckyballs, nanotubes, and stacked layers.

Crystal Defects

In a crystalline solid, the atoms, ions, or molecules are arranged in a definite repeating pattern, but occasional defects may occur in the pattern. Several types of defects are known, as illustrated in **Figure 1.28**. **Vacancies** are defects that occur when positions that should contain atoms or ions are vacant. Less commonly, some atoms or ions in a crystal may occupy positions, called **interstitial sites**, located between the regular positions for atoms. Other distortions are found in impure crystals, as, for example, when the cations, anions, or molecules of the impurity are too large to fit into the regular positions without distorting the structure. Trace amounts of impurities are sometimes added to a crystal (a process known as *doping*) in order to create defects in the structure that yield desirable changes in its properties. For example, silicon crystals are doped with varying amounts of different elements to yield suitable electrical properties for their use in the manufacture of semiconductors and computer chips.

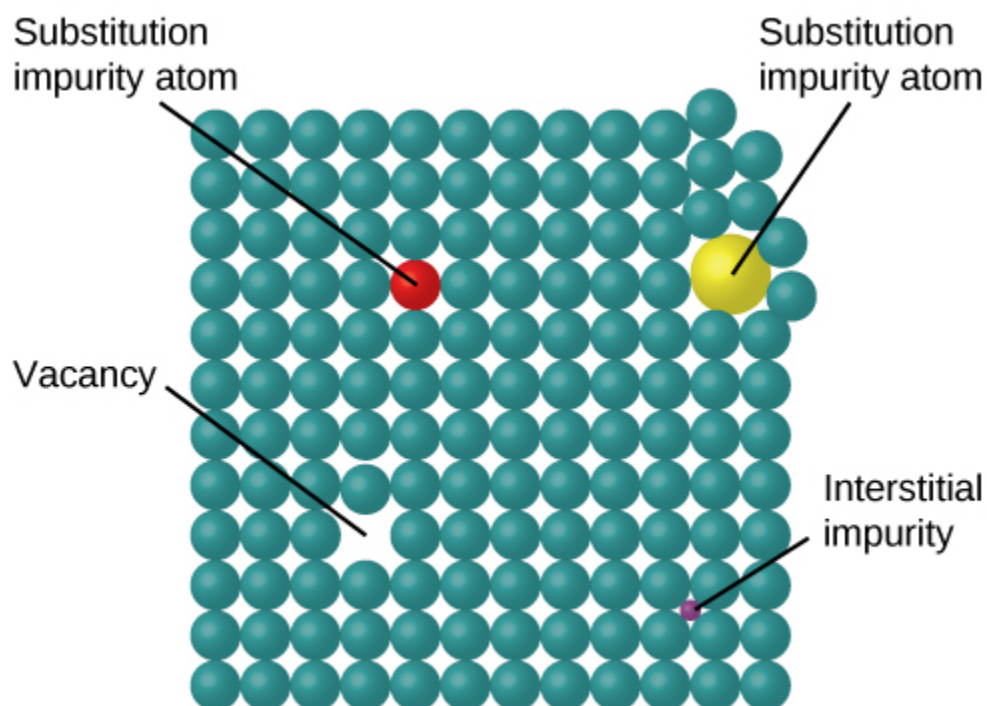


Figure 1.28 Types of crystal defects include vacancies, interstitial atoms, and substitutions impurities.

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 10. 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.

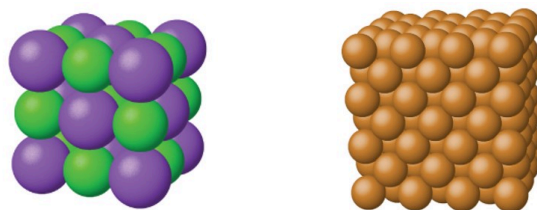


Figure 1.29 Both table salt (NaCl) on the left and solid copper on the right are simple crystalline solids with a cube-based structure.

In the Einstein solid model, we model the bonds between each atom as tiny springs that obey the Hooke's Law $F=kx$ that we discussed in Unit 2. This model is shown in Figure 11. 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.

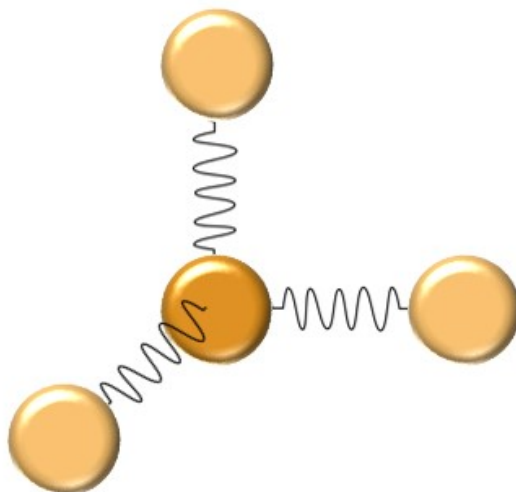


Figure 1.30 The model of an Einstein solid, the central atom is connected to its neighbors by atomic bonds that are modeled by springs that obey Hooke's Law

Key Concepts and Summary

Some substances form crystalline solids consisting of particles in a very organized structure; others form amorphous (noncrystalline) solids with an internal structure that is not ordered. The main types of crystalline solids are ionic solids, metallic solids, covalent network solids, and molecular solids. The properties of the different kinds of crystalline solids are due to the types of particles of which they consist, the arrangements of the particles, and the strengths of the attractions between them. Because their particles experience identical attractions, crystalline solids have distinct melting temperatures; the particles in amorphous solids experience a range of interactions, so they soften gradually and melt over a range of temperatures. Some crystalline solids have defects in the definite repeating pattern of their particles. These defects (which include vacancies, atoms or ions not in the regular positions, and impurities) change physical properties such as electrical conductivity, which is exploited in the silicon crystals used to manufacture computer chips.

Chemistry End of Chapter Exercises

Exercise 1.69

What types of liquids typically form amorphous solids?

Exercise 1.70

At very low temperatures oxygen, O_2 , freezes and forms a crystalline solid. Which best describes these crystals?

- (a) ionic
- (b) covalent network
- (c) metallic
- (d) amorphous
- (e) molecular crystals

Solution

- (e) molecular crystals

Exercise 1.71

As it cools, olive oil slowly solidifies and forms a solid over a range of temperatures. Which best describes the solid?

- (a) ionic
- (b) covalent network
- (c) metallic
- (d) amorphous
- (e) molecular crystals

Exercise 1.72

Explain why ice, which is a crystalline solid, has a melting temperature of 0 °C, whereas butter, which is an amorphous solid, softens over a range of temperatures.

Solution

Ice has a crystalline structure stabilized by hydrogen bonding. These intermolecular forces are of comparable strength and thus require the same amount of energy to overcome. As a result, ice melts at a single temperature and not over a range of temperatures. The various, very large molecules that compose butter experience varied van der Waals attractions of various strengths that are overcome at various temperatures, and so the melting process occurs over a wide temperature range.

Exercise 1.73

Identify the type of crystalline solid (metallic, network covalent, ionic, or molecular) formed by each of the following substances:

- (a) SiO₂
- (b) KCl
- (c) Cu
- (d) CO₂
- (e) C (diamond)
- (f) BaSO₄
- (g) NH₃
- (h) NH₄F
- (i) C₂H₅OH

Exercise 1.74

Identify the type of crystalline solid (metallic, network covalent, ionic, or molecular) formed by each of the following substances:

- (a) CaCl₂
- (b) SiC
- (c) N₂
- (d) Fe
- (e) C (graphite)
- (f) CH₃CH₂CH₂CH₃
- (g) HCl
- (h) NH₄NO₃
- (i) K₃PO₄

Solution

(a) ionic; (b) covalent network; (c) molecular; (d) metallic; (e) covalent network; (f) molecular; (g) molecular; (h) ionic; (i) ionic

Exercise 1.75

Classify each substance in the table as either a metallic, ionic, molecular, or covalent network solid:

Table 1.6

Substance	Appearance	Melting Point	Electrical Conductivity	Solubility in Water
X	lustrous, malleable	1500 °C	high	insoluble
Y	soft, yellow	113 °C	none	insoluble
Z	hard, white	800 °C	only if melted/dissolved	soluble

Exercise 1.76

Classify each substance in the table as either a metallic, ionic, molecular, or covalent network solid:

Table 1.7

Substance	Appearance	Melting Point	Electrical Conductivity	Solubility in Water
X	brittle, white	800 °C	only if melted/dissolved	soluble
Y	shiny, malleable	1100 °C	high	insoluble
Z	hard, colorless	3550 °C	none	insoluble

Solution

X = ionic; Y = metallic; Z = covalent network

Exercise 1.77

Identify the following substances as ionic, metallic, covalent network, or molecular solids:

Substance A is malleable, ductile, conducts electricity well, and has a melting point of 1135 °C. Substance B is brittle, does not conduct electricity as a solid but does when molten, and has a melting point of 2072 °C. Substance C is very hard, does not conduct electricity, and has a melting point of 3440 °C. Substance D is soft, does not conduct electricity, and has a melting point of 185 °C.

Exercise 1.78

Substance A is shiny, conducts electricity well, and melts at 975 °C. Substance A is likely a(n):

- (a) ionic solid
- (b) metallic solid
- (c) molecular solid
- (d) covalent network solid

Solution

(b) metallic solid

Exercise 1.79

Substance B is hard, does not conduct electricity, and melts at 1200 °C. Substance B is likely a(n):

- (a) ionic solid
- (b) metallic solid
- (c) molecular solid
- (d) covalent network solid

Glossary

amorphous solid: (also, noncrystalline solid) solid in which the particles lack an ordered internal structure

anion: negatively charged atom or molecule (contains more electrons than protons)

atomic mass: average mass of atoms of an element, expressed in amu

atomic mass unit (amu): (also, unified atomic mass unit, u, or Dalton, Da) unit of mass equal to $\frac{1}{12}$ of the mass of a ^{12}C atom

atomic number (Z): number of protons in the nucleus of an atom

Avogadro's number: N_A , the number of molecules or atoms in one mole of a substance; $N_A = 6.02 \times 10^{23}$ particles/mole

Avogadro's number (N_A): experimentally determined value of the number of entities comprising 1 mole of substance, equal

to $6.022 \times 10^{23} \text{ mol}^{-1}$

Boltzmann constant: k , a physical constant that relates energy to temperature; $k = 1.38 \times 10^{-23} \text{ J/K}$

cation: positively charged atom or molecule (contains fewer electrons than protons)

chemical symbol: one-, two-, or three-letter abbreviation used to represent an element or its atoms

covalent network solid: solid whose particles are held together by covalent bonds

crystalline solid: solid in which the particles are arranged in a definite repeating pattern

Dalton (Da): alternative unit equivalent to the atomic mass unit

formula mass: sum of the average masses for all atoms represented in a chemical formula; for covalent compounds, this is also the molecular mass

fundamental unit of charge: (also called the elementary charge) equals the magnitude of the charge of an electron (e) with $e = 1.602 \times 10^{-19} \text{ C}$

ideal gas law: the physical law that relates the pressure and volume of a gas to the number of gas molecules or number of moles of gas and the temperature of the gas

interstitial sites: spaces between the regular particle positions in any array of atoms or ions

ion: electrically charged atom or molecule (contains unequal numbers of protons and electrons)

ionic solid: solid composed of positive and negative ions held together by strong electrostatic attractions

mass number (A): sum of the numbers of neutrons and protons in the nucleus of an atom

metallic solid: solid composed of metal atoms

molar mass: mass in grams of 1 mole of a substance

mole: amount of substance containing the same number of atoms, molecules, ions, or other entities as the number of atoms in exactly 12 grams of ^{12}C

mole: the quantity of a substance whose mass (in grams) is equal to its molecular mass

molecular solid: solid composed of neutral molecules held together by intermolecular forces of attraction

unified atomic mass unit (u): alternative unit equivalent to the atomic mass unit

vacancy: defect that occurs when a position that should contain an atom or ion is vacant

2 INTRODUCTION TO ENERGY

2.1 Introduction to Energy

Up to this point in the course, we've really been talking about forces. Now we're going to change things up and talk about a quantity that you may be familiar with from previous science courses: the idea of energy. This chapter introduces this idea. In the next chapter, we'll begin our formal study of energy by looking at some of the same macroscopic phenomena we have been discussing already: falling objects, sliding with friction, and other such situations. We'll just be looking at these same situations from a different viewpoint. It turns out that you can describe any phenomenon we've talked about in the course up to this point either way, either with energy or with forces, although often, one is easier than the other. In the subsequent chapter we'll move on and use our idea of energy at the microscopic scale to understand completely new phenomena that are very directly related to ideas you might have discussed in the chemistry or a biology course.

So, what is energy? in short, energy is the ability to do work. That's the definition that we'll be using in this course. You can also flip this on its head with the change of perspective, however, and say that work is one way of transferring this quantity known as energy. Before we get too deep into it, it's probably worth comparing briefly the idea of energy and the idea forces. So, up to this point we've been talking about forces, and forces are an idea of instance. We look at the forces that are acting on an object "right now", and "right now" is all that matters, and then we can work instant by instant like we did in our simulations to study the motion of an object. Most physics courses begin with the idea of forces because they're easy to get a feel for what a force is., we've all experienced pushes and pulls.

Energy, on the other hand, is an example of a conservation law. Energy is a quantity that never changes through a process. This allows us to relate two points in time that are not directly next to each other. So, with forces, we had to take very tiny little time steps. With energy, we can go from the beginning of a process all the way to the end, not care too much about the middle, and know that energy is going to be conserved. So, why don't we start with energy? Well, energy can be conceptually a little bit more difficult to get your head around. What exactly is energy; it's a little bit more abstract of an idea, which is why most physics courses, including this one, put it off until after a discussion of forces.

So, what's the big point of the next three chapters? So, you've probably dealt with energy before in a previous science class. Our goal is to develop a coherent picture of energy across different scientific disciplines and across a large variety of different distant scales, from the sizes that we experience in our everyday world of people and cars and trees, all the way down to the atomic and molecular scale. We will, therefore, deal with several different ideas that can seem unrelated. We'll talk about boxes on hills and fish on springs and the kinetic energy of, say, cars, but we'll also talk about the kinetic energy of molecules, which we'll find out is directly related to the idea of temperature, and we'll talk about energy transfer through random collisions of particles on the atomic scale, which we'll talk about as the idea of heat. These different ideas and others including chemical bonds are connected by the idea of energy.

So, again, let's review how the next three chapters are laid out this current chapter provides a basic overview of what energy is, identifies the two main types of energy, kinetic and potential energy, identifies two main scales of energy we'll discuss in this class, the macroscopic scale of people, cars, etc., and the microscopic scale of atoms and molecules. The next chapter discusses energy on this macroscopic scale of people cars and so forth, and the chapter after that really gets into energy on the microscopic scale, atoms, molecules, and temperature.

2.2 Units of Energy

UMASS AMHERST Instructor's Notes

Your Quiz will Cover

- Converting between the different 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 $1\text{J}=1\text{Nm}=1\text{kgms}^2$). 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.

UMASS AMHERST Instructor's Notes

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, $1\text{eV} = 1.602 \times 10^{-19}\text{J}$. 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]. The bond-dissociation energy is the energy released when the bond is formed.

Table 2.1

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
H ₃ C-H	4.550	One of the strongest aliphatic C–H bonds

UMASS AMHERST Instructor's Notes

Note that the eV is significantly smaller than the joule; eV will generally be the smallest unit of energy used in this course.

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: $1\text{W} = 1\text{J/s}$. Thus, a kilowatt hour is:

$$(1\text{kW} * \text{hr}) \left(\frac{1000\text{W}}{\text{kW}} \right) \left(\frac{1\text{J/s}}{\text{W}} \right) \left(\frac{3600\text{s}}{\text{hr}} \right) = 3.6106\text{J} = 3.6\text{MJ}$$

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 1°C or $1\text{cal} = 4.184\text{J}$. On food labels, you will see energy listed in Calories (capital C). One Calorie is equal to 1kilocalorie; in other words, $1\text{Cal} = 1000\text{cal}$. 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 1.



Figure 2.1 A food label from the UK showing the energy of the food in both Joules and kcal (or Calories).

1. "Bond-Dissociation Energy - Wikipedia." Accessed August 1, 2017. https://en.wikipedia.org/wiki/Bond-dissociation_energy.

2.3 Types of Energy and Scales of Energy

UMASS AMHERST Instructor's Notes

This section provides the definitions of the terms kinetic and potential energy and defines the microscopic and macroscopic scales. Details and formulae will be provided in upcoming chapters.

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 2 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 2). The sheer fact that the car is moving means that it *can* do work. Similarly, the larger block in Figure 3 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 2.2 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.^[2]

2. 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>.

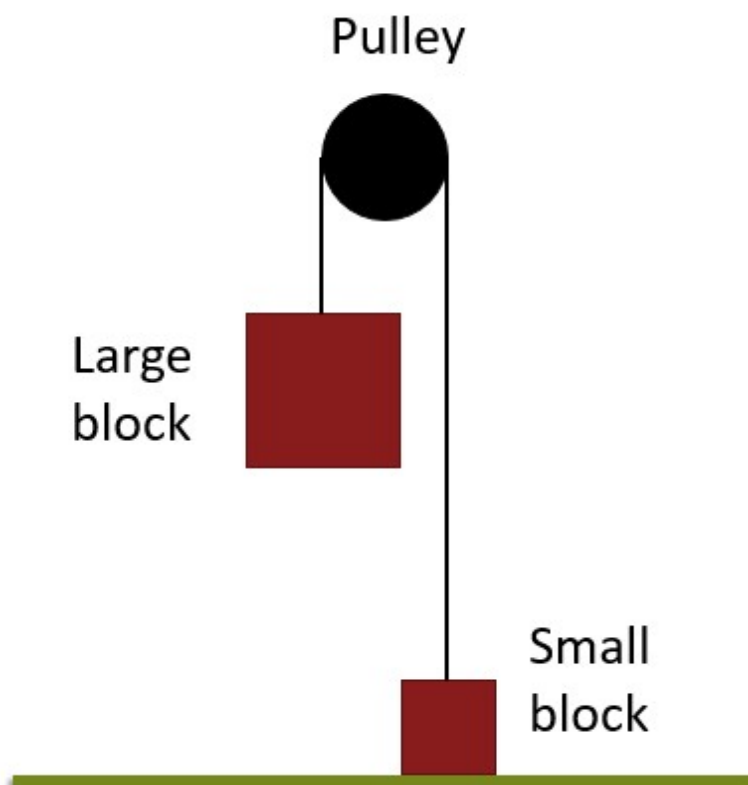


Figure 2.3 A large block connected to a small block over a pulley has an ability to do work due to its position relative to the earth; the large block has potential energy. We see that work when the large block is released exerting, via the rope and pulley, a force on the small block for a distance causing 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.

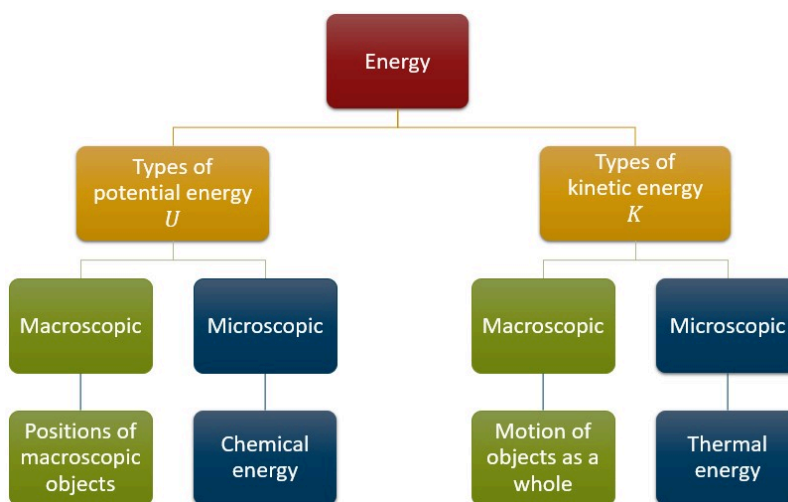


Figure 2.4 The relationships between different types of energy

UMASS AMHERST Instructor's Notes

This chart provides a way to organize the different types of energy. We will flush out this chart in upcoming sections.

2.4 Conservation of Energy

UMASS AMHERST Instructor's Notes

Your Quiz will Cover

- Understanding that, in a closed system, energy is conserved

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: $E_i + E_{in} - E_{out} = E_f$.

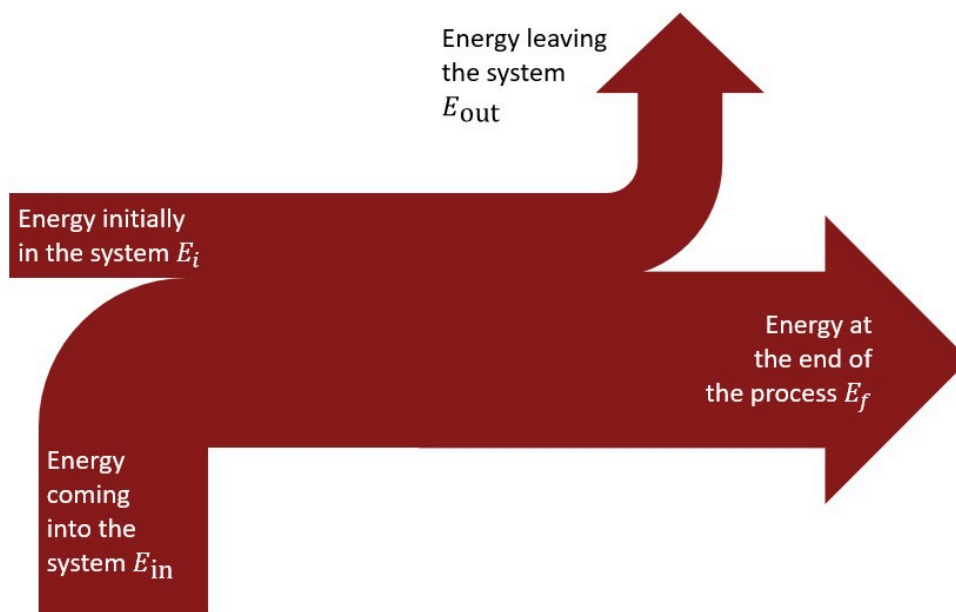


Figure 2.5 The conservation of energy in graphical form $E_i + E_{in} - E_{out} = E_f$

2.5 Ways to Transfer Energy

UMASS AMHERST Instructor's Notes

Your Quiz will Cover

- Understanding the difference between heat and temperature
- Understanding the differences between the different types of heat transfer

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 Q . At its core, *heat is the transfer of energy by collisions at the microscopic scale.*

Heat

We already know that work is one way in which we can transfer energy, but you might be thinking to yourself, is there another way we can transfer energy into or out of a system? We know, for example, that if I place a warm soda in contact with the cold ice cube, the soda cools and some of the ice cube melts until the two come to a constant temperature. It certainly seems that energy is being transferred in this process; how is the energy being transferring in a process like this?

This method of energy transfer is known as heat. Heat is a transfer of energy; an object cannot have heat any more than an object can have work. It's also important to keep in mind the distinction between heat and temperature. The symbol for heat that we will use is the letter Q . Since it is a measure of transferred energy, heat will have the unit of joules, just like work and energy do. The sign convention for heat that we will adopt is that when energy flows into a system, Q is positive, and when energy flows out of a system, heat is negative.

There are three main ways to transfer heat into or out of a system. The first method is conduction, which is essentially when you place a metal stick in a fire, the other end of the stick gets hot, the second method is convection, which occurs when you have say a pot of water over a flame, and the third method is radiation, where light from the Sun warms the earth. Let's now explore these three processes in a little bit more detail.

Conduction

Conduction occurs when atoms in a material collide with each other. So, let's assume we have some sort of solid, and of course inside the solid the atoms are arranged in some regular pattern. When we place one end the solid in the flame, the atoms nearest the flame gain energy and begin to vibrate more vigorously, increasing their temperature. These atoms bump into the atoms next to them, which bump into the atoms next for them, and so on and so on, transferring energy up the rod. This transfer of energy through the molecular collisions of the atoms within the rod is known as heat. In a solid, the atoms are very close together and can easily bump into each other, and therefore solids transfer heat quite well. In liquids and gases, on the other hand, the atoms are a little further apart, and so don't collide as often. They can still transfer energy as heat through conduction, just not as well, because the collisions won't be as frequent. In a metal on the other hand, the e^- electrons that bind the metal together are free to move around and collide with each other. These e^- electrons, since they can travel so far, result in a lot of collisions, and result in metals conducting heat very well. Heat will quickly result in the transfer of energy from one end of a piece of metal to the other, which is why it's easy to burn your hand on a piece of metal that's hot on one end.

Radiation

So, in our first example, the fire caused the atoms nearest to it in the solid to begin to move. Well, how did this happen? Well the fire generates particles of light energy called photons that will be discussed in some detail in Physics 132. These photons, some of them travel and collide with the atoms in the solid, giving them their energy and causing them to move, and setting off the chain that we talked about in conduction. This transfer of energy through molecular collisions involving photons is called heat by radiation. This is the mechanism by which the earth gets its energy from the Sun. There is no matter in between the earth and the Sun to provide a mechanism for conduction. The only mechanism of heat transfer is through radiation of photons emitted by the Sun, which can travel through the vacuum of space.

Convection

The final way the transfer heat is called convection. Now, convection is a little bit more complicated than the other mechanisms, so we won't go into it in detail, but in short, convection is the transfer of energy through the bulk motion of a fluid, so a fluid in physics means either a liquid or a gas, and by the bulk motion, we mean giant currents of liquid or gas moving through the material. In our example of a pot of water on a stove, the liquid near the bottom of the pot gains energy through conduction. Flame gets the bottom of the pot warm through radiation, and the bottom of the pot is in contact with the liquid and transfers energy by conduction. Now the water at the bottom of the pot is warmer than the surrounding. Warmer fluids tend to have lower density, so this warm fluid near the bottom rises to the top carrying its energy with it, and cooler liquid settles to the bottom, and thus we get this current up in the middle and down the outside, bulk motion of the water transferring energy through the system. This is something that we see a lot and is of critical importance in the motion of the oceans. Cool water from the poles sinks to the bottom of the ocean, travels down to the equator, where it's warmed up by the Sun, and then rises back to the top, distributing energy and nutrients throughout the ocean.

Summary

So, now we've talked about the three methods of heat transfer, conduction, radiation, and convection. What's the commonality between these three different methods? These three different methods are, at the microscopic level, the transfer of energy through random collisions. In conduction, the atoms are colliding with each other, in radiation you have the collision of photons with atoms, and in convection you have the transfer of energy through the motion of the fluid itself. And this is what heat really is, it's the transfer of energy through collisions at the microscopic scale.

Heat, which we will use the letter Q to represent, is another way of transferring energy. Energy entering the system has positive heat in our convention, and energy leaving the system has negative heat in our convention. This energy transfer can occur in one of three ways, conduction, convection, or radiation, but these methods are, at the molecular level, collisions between particles.

UMASS AMHERST Instructor's Notes

(insert text here)

This paragraph summarizes things nicely and is important for your quiz!

You need to know what heat is and how it is different from temperature. You also need to know that if energy is going into a system, then we count it as positive heat or work. Energy leaving is negative heat or work.

2.6 The Formal Statement of the Conservation of Energy as the First Law of Thermodynamics

UMASS AMHERST Instructor's Notes

While this section is more formally mathematical than most, you should pay attention as this is probably the most important idea in all of science! For your quiz, you will need to be able to use this idea in the abstract. If I give you two of ΔE , Q , or W you should be able to tell me the third.

In section 13.4, we stated that energy must be conserved: $E_i + E_{in} - E_{out} = E_f$. Moving things around we get $E_f - E_i = E_{in} - E_{out}$. Recognizing the term on the left as ΔE , we can say $\Delta E = E_{in} - E_{out}$. If we redefine E_{in} and E_{out} as just different directions of transferred energy $E_{transferred}$, then we have $\Delta E = E_{transferred}$ where $E_{transferred}$ 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 Q and work W . Thus, $E_{transferred}$ must be the sum of the energy transferred by heat and the energy transferred by work, $E_{transferred} = Q + W$. The statement of the law of conservation of energy can therefore be written as

$$\Delta E = Q + W$$

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

$$E_f - E_i = Q + W$$

where both heat Q and work W 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, $W > 0$ as energy is coming in and $Q = 0$. The result is that $E_f > E_i$, 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 $Q < 0$ because heat is flowing out and $W = 0$ due to the fact that there is no “distance” for $W = Fd\cos\theta$. Therefore, $Q + W < 0$ and $E_f < E_i$ as expected given that energy is flowing out of the system.

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

UMASS AMHERST Instructor's Notes

There are some differences in conventions when expressing the first law of thermodynamics to be aware of.

In some other courses or references, you may see the first law of thermodynamics written as $\Delta E = Q - W$, 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 Q as well as matching our definition of work from mechanics $W = Fd\cos\theta$ which considers only external forces. Physically, we are thinking about work

done *on* the system by *external* forces.

To understand the $\Delta E = Q - W$ 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 $\vec{F}_{\text{environment} \rightarrow \text{engine}}$, they were thinking about $\vec{F}_{\text{engine} \rightarrow \text{environment}}$. 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 $\Delta E = Q - W$ instead of $\Delta E = Q + W$.

In this class, we will stick with $\Delta E = Q + W$, 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 $\Delta E = Q - W$ 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?

2.8 Enthalpy

UMASS AMHERST Instructor's Notes

For your quiz, you will need to know the distinctions between enthalpy, energy, heat, and work

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 (ΔH) is:

$$\Delta H = \Delta U + P\Delta V$$

Recall from the chapter on work that the mathematical product $P\Delta V$ represents work (W). If I compress the gas in Figure 1, 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 ΔV and w will always be opposite:

$$P\Delta V = -W$$

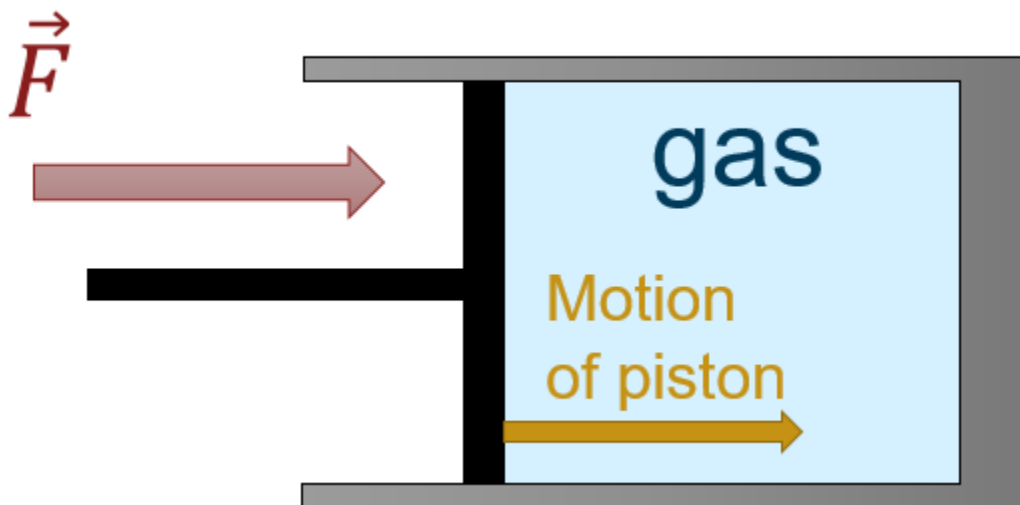


Figure 2.6 A gas being compressed by a piston. The force on the piston is in the same direction as the piston's motion implying positive work. However, the volume of the gas is shrinking meaning $\Delta V < 0$

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

$$\Delta H = \Delta E + P\Delta V$$

$$\Delta H = (QP + W) - W$$

$$\Delta H = QP$$

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 (ΔH) 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 ΔH 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 $Q = \Delta H$.

3 ENERGY OF OBJECTS AS A WHOLE (MACROSCOPIC SCALE)

3.1 Introduction

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.*

3.2 Kinetic Energy of an Object

UMASS AMHERST Instructor's Notes

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.

Our goal in this section is to figure out an expression for the kinetic energy.

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 1.

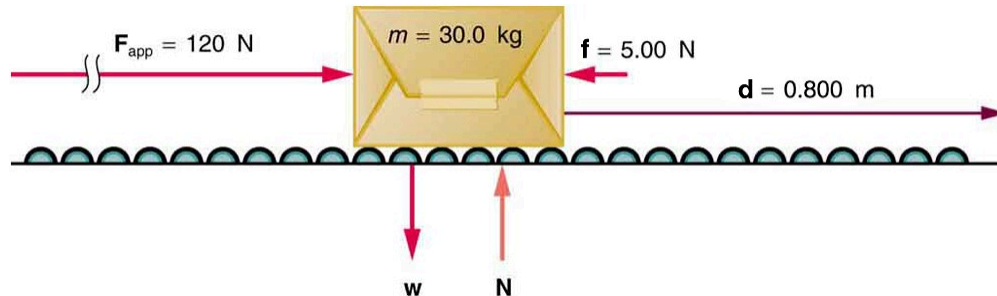


Figure 3.1 A package on a roller belt is pushed horizontally through a distance d

In this case, there is no transfer of energy by molecular collisions, i.e. there is no heat and $Q=0$. Meaning that our statement of conservation of energy goes from

$$E = Q + W$$

to

$$E = W$$

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: $E=K$. Therefore, our statement of conservation of energy for this situation is just

$$\Delta K = W$$

$$K_f - K_i = W$$

The effect of the net force F_{net} 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 (https://cnx.org/contents/Ax2o07UI@9.86:P_-6tVsN@5/Kinetic-Energy-and-the-Work-En#fs-id1703845) Example (https://cnx.org/contents/Ax2o07UI@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 $\theta=0^\circ$ and $\cos\theta=1$, 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 (<https://cnx.org/contents/031da8d3-b525-429c-80cf-6c8ed997733a@9.86:ea2bb23c-4fce-4e9d-a46b-3754125da988@9>) **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, $v^2 = v_0^2 + 2ad$ (note that a appears in the expression for the net work). Solving for acceleration gives $a = (v^2 - v_0^2)/2d$. When a is substituted into the preceding expression for W_{net} , we obtain

$$W_{net} = m\left(\frac{v^2 - v_0^2}{2d}\right)d$$

The d cancels, and we rearrange this to obtain

$$W_{net} = \frac{1}{2}mv^2 - \frac{1}{2}mv_0^2$$

UMASS AMHERST Instructor's Notes

The following subsection is based on umdberg / Kinetic energy and the work-energy theorem (2013). Available at: [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)). ([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))) (Accessed: 9th August 2017)

Interpreting the Result: Kinetic Energy

UMASS AMHERST Instructor's Notes

While, this section is short, it is arguably more important than the text above. This section is about interpreting our expression for kinetic energy and helping you understand how it is different from the, at first glance, similar looking momentum.

This section is based upon [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)) ([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)))

What has come out after all our manipulations is that the work in this case is related to a change in a quantity associated with motion, $\frac{1}{2}mv^2$. 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 $\frac{1}{2}mv^2$ must be the *kinetic energy*. It is a measure of "the energy associated with how much an object is moving".

3.3 Examples Applying Conservation of Energy with only Kinetic Energy

UMASS AMHERST Instructor's Notes

For your quiz, you are expected to be able to solve problems such as these.

Example 3.1 Calculating the Kinetic Energy of a Package

Suppose a 30.0-kg package on the roller belt conveyor system in ??? is moving at 0.500 m/s. What is its kinetic energy?

Strategy

Because the mass m and speed v are given, the kinetic energy can be calculated from its definition as given in the equation $KE = \frac{1}{2}mv^2$.

Solution

The kinetic energy is given by

$$KE = \frac{1}{2}mv^2. \quad (3.1)$$

Entering known values gives

$$KE = 0.5(30.0 \text{ kg})(0.500 \text{ m/s})^2, \quad (3.2)$$

which yields

$$KE = 3.75 \text{ kg} \cdot \text{m}^2/\text{s}^2 = 3.75 \text{ J}. \quad (3.3)$$

Discussion

Note that the unit of kinetic energy is the joule, the same as the unit of work, as mentioned when work was first defined. It is also interesting that, although this is a fairly massive package, its kinetic energy is not large at this relatively low speed. This fact is consistent with the observation that people can move packages like this without exhausting themselves.

Example 3.2 Kinetic Energy of a Car

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

$$\frac{K_f}{K_i}$$

Substituting our definition for kinetic energy we get

$$\frac{K_f}{K_i} = \frac{\frac{1}{2}mv_f^2}{\frac{1}{2}mv_i^2}$$

Since the mass of the car does not change, m and the $\frac{1}{2}$ cancel leaving

$$\frac{K_f}{K_i} = \frac{v_f^2}{v_i^2}$$

Substituting our values in we get

$$\frac{K_f}{K_i} = \frac{10^2}{5^2} = \frac{100}{25} = 4$$

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!

Example 3.3 Deep Space 1

Deep Space 1 was a space probe, launched on October 24, 1998, and it used a type of engine called a ion propulsion drive. This engine generates a weak force, but it can do so over a long period of time and using only a small amount of fuel. The probe has a mass of 474 kg and is traveling with an initial speed of 275 m/s. The only force acting on the probe is from the ion drive, at a force of 0.056N parallel to the probe's displacement, which is 2.42 million km. What is the final speed of the probe?

Solution

In this problem, we're looking for a change in speed, which tells us that we're looking for a change in the kinetic energy. With the only force on the probe being the ion drive, we can use the work-energy theorem to find the change in energy:

$$W = \Delta E$$

Then, we can expand the energy: $W = E_f - E_i$

Since the only energy in this system is the kinetic energy, we can substitute the energies with our definition of kinetic energy:

$$W = \frac{1}{2}mv_f^2 - \frac{1}{2}mv_i^2$$

Since it's the final velocity that we're looking for, we can solve this equation for the final velocity:

$$v_f = \sqrt{\frac{W + \frac{1}{2}mv_i^2}{\frac{1}{2}m}} = \sqrt{\frac{2W}{m} + v_i^2}$$

The mass and the initial velocity are both given, but we need to find the work done on the probe. We can use our definition of work to calculate it:

$$W = F\Delta x = 0.056N * 2.42 \times 10^9 m = 135520000J$$

Substituting our values and calculating the final velocity: $v_f = \sqrt{\frac{2(135520000J)}{474kg} + (275m/s)^2} = 804m/s$

3.4 Macroscopic Potential Energy**Work Done Against Gravity**

Climbing stairs and lifting objects is work in both the scientific and everyday sense—it is work done against the gravitational force. When there is work, there is a transformation of energy. The work done against the gravitational force goes into an important form of stored energy that we will explore in this section.

Let us calculate the work done in lifting an object of mass m through a height h , such as in **Figure 3.2**. If the object is lifted straight up at constant speed, then the force needed to lift it is equal to its weight mg . The work done on the mass is then

$W = Fd = mgh$. We define this to be the **gravitational potential energy** (U_g) put into (or gained by) the object-Earth

system. This energy is associated with the state of separation between two objects that attract each other by the gravitational force. For convenience, we refer to this as the U_g gained by the object, recognizing that this is energy stored in the gravitational field of Earth. Why do we use the word "system"? Potential energy is a property of a system rather than of a single object—due to its physical position. An object's gravitational potential is due to its position relative to the surroundings within the Earth-object system. The force applied to the object is an external force, from outside the system. When it does positive work it increases the gravitational potential energy of the system. Because gravitational potential energy depends on relative position, we need a reference level at which to set the potential energy equal to 0. We usually choose this point to be Earth's surface, but this point is arbitrary; what is important is the *difference* in gravitational potential energy, because this difference is what relates to the work done. The difference in gravitational potential energy of an object (in the Earth-object system) between two rungs of a ladder will be the same for the first two rungs as for the last two rungs.

Converting Between Potential Energy and Kinetic Energy

Gravitational potential energy may be converted to other forms of energy, such as kinetic energy. If we release the mass, gravitational force will do an amount of work equal to mgh on it, thereby increasing its kinetic energy by that same amount (by

the work-energy theorem). We will find it more useful to consider just the conversion of U_g to KE without explicitly considering the intermediate step of work. (See ???.) This shortcut makes it easier to solve problems using energy (if possible) rather than explicitly using forces.

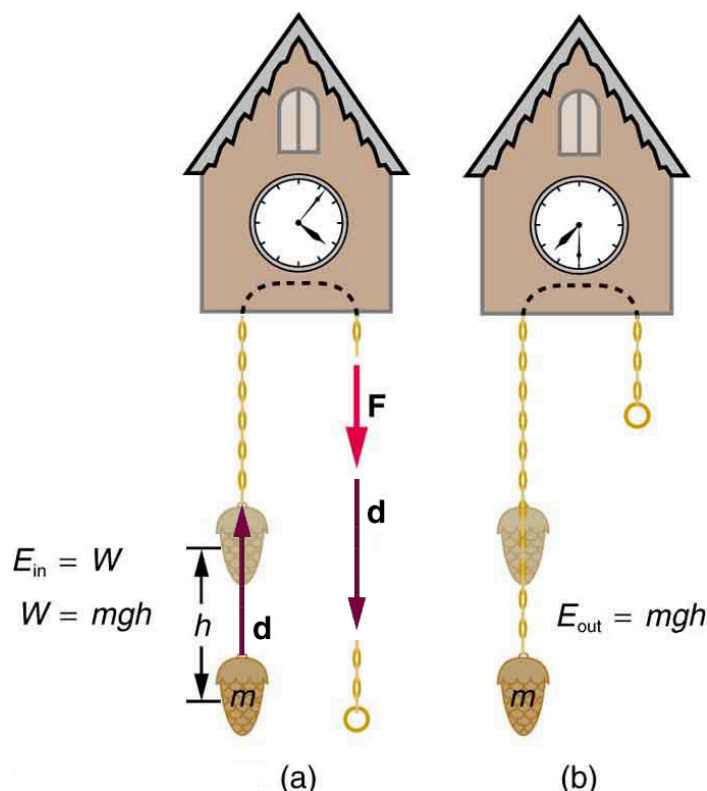


Figure 3.2 (a) The work done to lift the weight is stored in the mass-Earth system as gravitational potential energy. (b) As the weight moves downward, this gravitational potential energy is transferred to the cuckoo clock.

More precisely, we define the *change* in gravitational potential energy ΔU_g to be

$$\Delta U_g = mgh, \quad (3.4)$$

where, for simplicity, we denote the change in height by h rather than the usual Δh . Note that h is positive when the final height is greater than the initial height, and vice versa. For example, if a 0.500-kg mass hung from a cuckoo clock is raised 1.00 m, then its change in gravitational potential energy is

$$\begin{aligned} mgh &= (0.500 \text{ kg})(9.80 \text{ m/s}^2)(1.00 \text{ m}) \\ &= 4.90 \text{ kg} \cdot \text{m}^2/\text{s}^2 = 4.90 \text{ J}. \end{aligned} \quad (3.5)$$

Note that the units of gravitational potential energy turn out to be joules, the same as for work and other forms of energy. As the clock runs, the mass is lowered. We can think of the mass as gradually giving up its 4.90 J of gravitational potential energy, *without directly considering the force of gravity that does the work*.

Using Potential Energy to Simplify Calculations

The equation $\Delta U_g = mgh$ applies for any path that has a change in height of h , not just when the mass is lifted straight up.

(See **Figure 3.3**.) It is much easier to calculate mgh (a simple multiplication) than it is to calculate the work done along a complicated path. The idea of gravitational potential energy has the double advantage that it is very broadly applicable and it makes calculations easier. From now on, we will consider that any change in vertical position h of a mass m is accompanied by a change in gravitational potential energy mgh , and we will avoid the equivalent but more difficult task of calculating work done by or against the gravitational force.

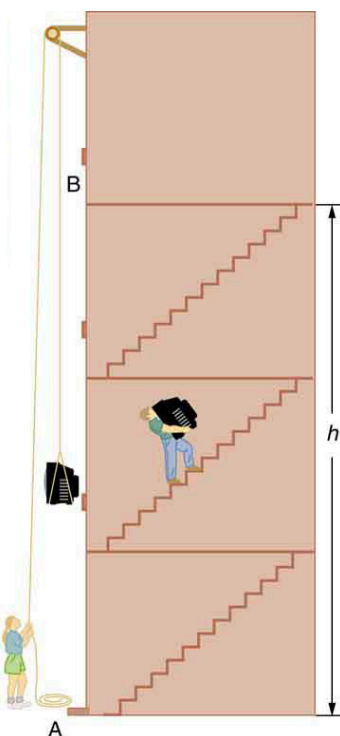


Figure 3.3 The change in gravitational potential energy (ΔU_g) between points A and B is independent of the path. $\Delta U_g = mgh$ for any path between the two points. Gravity is one of a small class of forces where the work done by or against the force depends only on the starting and ending points, not on the path between them.

Example 3.4 The Force to Stop Falling

A 60.0-kg person jumps onto the floor from a height of 3.00 m. If he lands stiffly (with his knee joints compressing by 0.500 cm), calculate the force on the knee joints.

Strategy

This person's energy is brought to zero in this situation by the work done on him by the floor as he stops. The initial U_g is transformed into KE as he falls. The work done by the floor reduces this kinetic energy to zero.

Solution

The work done on the person by the floor as he stops is given by

$$W = Fd \cos \theta = -Fd, \quad (3.6)$$

with a minus sign because the displacement while stopping and the force from floor are in opposite directions ($\cos \theta = \cos 180^\circ = -1$). The floor removes energy from the system, so it does negative work.

The kinetic energy the person has upon reaching the floor is the amount of potential energy lost by falling through height h :

$$KE = -\Delta U_g = -mgh, \quad (3.7)$$

The distance d that the person's knees bend is much smaller than the height h of the fall, so the additional change in gravitational potential energy during the knee bend is ignored.

The work W done by the floor on the person stops the person and brings the person's kinetic energy to zero:

$$W = -KE = mgh. \quad (3.8)$$

Combining this equation with the expression for W gives

$$-Fd = mgh. \quad (3.9)$$

Recalling that h is negative because the person fell *down*, the force on the knee joints is given by

$$F = -\frac{mgh}{d} = -\frac{(60.0 \text{ kg})(9.80 \text{ m/s}^2)(-3.00 \text{ m})}{5.00 \times 10^{-3} \text{ m}} = 3.53 \times 10^5 \text{ N}. \quad (3.10)$$

Discussion

Such a large force (500 times more than the person's weight) over the short impact time is enough to break bones. A much better way to cushion the shock is by bending the legs or rolling on the ground, increasing the time over which the force acts. A bending motion of 0.5 m this way yields a force 100 times smaller than in the example. A kangaroo's hopping shows this method in action. The kangaroo is the only large animal to use hopping for locomotion, but the shock in hopping is cushioned by the bending of its hind legs in each jump. (See [Figure 3.4](#).)



Figure 3.4 The work done by the ground upon the kangaroo reduces its kinetic energy to zero as it lands. However, by applying the force of the ground on the hind legs over a longer distance, the impact on the bones is reduced. (credit: Chris Samuel, Flickr)

We have seen that work done by or against the gravitational force depends only on the starting and ending points, and not on the path between, allowing us to define the simplifying concept of gravitational potential energy. We can do the same thing for a few other forces, and we will see that this leads to a formal definition of the law of conservation of energy.

The Zero of Potential Energy

This is perhaps a slightly subtler topic than you might first imagine. Let's think about this quantity of gravitational potential energy written as mgh . Here we have Mr. clumsy dropping a ball.



Figure 3.5 Mr. Clumsy dropping the ball.

Where is the gravitational potential energy of the ball equal to zero? Well, m and g are both numbers, m is the mass of the ball and g is 9.8 m/s^2 , so essentially this is the same question as where is height equal to 0, h . Well, the logical place you might think for h to be equal to 0 would be the ground. We define the ground to be $h=0$, then the gravitational potential energy on the ground, U_g , is going to be zero when the ball is on the ground. Now let's move Mr. Clumsy to a platform on the top of a skyscraper.

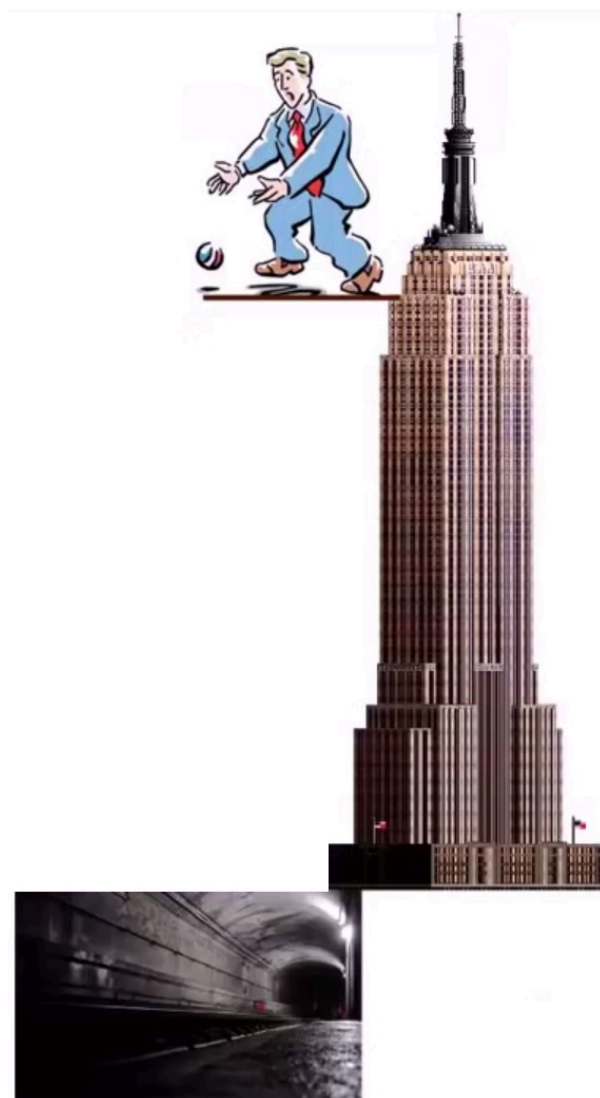


Figure 3.6 Mr. Clumsy dropping the ball on top of a skyscraper, with a subway tunnel below.

Now where is $h=0$? This is a perhaps a little bit trickier question; do we define $h=0$ to be at the platform, or do we define it to be in the ground, or do we even define it to be in the subway tunnel underneath the skyscraper? Which of these should we choose for $h=0$, or which of these should we choose for the zero of gravitational potential energy? Well, the universe doesn't care where we choose h to be equal to 0. so due to that, any of these choices, the platform, the ground, the subway tunnel, they're all fine, doesn't matter which we pick. You should just be very explicit with your choice, so when you're approaching a problem with gravitational potential energy, explicitly write down that I am going to choose the zero of gravitational potential energy to be the platform, for example. If I choose the ground, say, to have zero potential gravitational potential energy, then points in the subway tunnel below the ground have negative gravitational potential energy. There's absolutely nothing wrong with that.

Why is there nothing wrong with negative potential energy? Well, the work done by gravity is equal to the negative of the change in gravitational potential energy, or

$$W = -\Delta U$$

ΔU is always $U_f - U_i$, so the work done is

$$W = U_i - U_f$$

or using mgh for potential energy

$$W = mgh_i - mgh_f$$

The key point in all of this is that the work done does not depend upon the value of potential energy itself, only the change in the potential energy is a relevant quantity, and the change in potential.

Potential Energy of a Spring

First, let us obtain an expression for the potential energy stored in a spring (U_s). We calculate the work done to stretch or compress a spring that obeys Hooke's law. (Hooke's law was examined in **Elasticity: Stress and Strain** (<https://legacy.cnx.org/content/m42081/latest/>), and states that the magnitude of force F on the spring and the resulting deformation ΔL are proportional, $F = k\Delta L$.) (See **Figure 3.7**.) For our spring, we will replace ΔL (the amount of deformation produced by a force F) by the distance x that the spring is stretched or compressed along its length. So the force needed to stretch the spring has magnitude $F = kx$, where k is the spring's force constant. The force increases linearly from 0 at the start to kx in the fully stretched position. The average force is $kx/2$. Thus the work done in stretching or compressing the spring is $W_s = Fd = \left(\frac{kx}{2}\right)x = \frac{1}{2}kx^2$. Alternatively, we noted in **Kinetic Energy and the Work-Energy Theorem** (<https://legacy.cnx.org/content/m42147/latest/>) that the area under a graph of F vs. x is the work done by the force. In **Figure 3.7(c)** we see that this area is also $\frac{1}{2}kx^2$. We therefore define the **potential energy of a spring**, U_s , to be

$$U_s = \frac{1}{2}kx^2, \quad (3.11)$$

where k is the spring's force constant and x is the displacement from its undeformed position. The potential energy represents the work done *on* the spring and the energy stored in it as a result of stretching or compressing it a distance x . The potential energy of the spring U_s does not depend on the path taken; it depends only on the stretch or squeeze x in the final configuration.

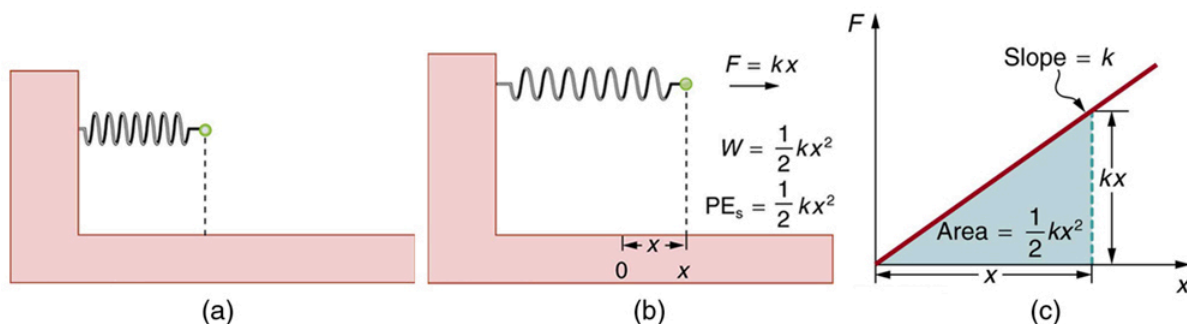


Figure 3.7 (a) An undeformed spring has no U_s stored in it. (b) The force needed to stretch (or compress) the spring a distance x has a magnitude $F = kx$, and the work done to stretch (or compress) it is $\frac{1}{2}kx^2$. Because the force is conservative, this work is stored as potential energy (U_s) in the spring, and it can be fully recovered. (c) A graph of F vs. x has a slope of k , and the area under the graph is $\frac{1}{2}kx^2$. Thus the work done or potential energy stored is $\frac{1}{2}kx^2$.

The equation $U_s = \frac{1}{2}kx^2$ has general validity beyond the special case for which it was derived. Potential energy can be stored in any elastic medium by deforming it. Indeed, the general definition of **potential energy** is energy due to position, shape, or configuration. For shape or position deformations, stored energy is $U_s = \frac{1}{2}kx^2$, where k is the force constant of the particular system and x is its deformation. Another example is seen in **Figure 3.8** for a guitar string.

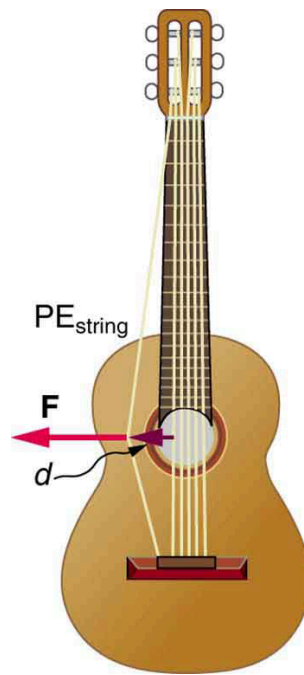


Figure 3.8 Work is done to deform the guitar string, giving it potential energy. When released, the potential energy is converted to kinetic energy and back to potential as the string oscillates back and forth. A very small fraction is dissipated as sound energy, slowly removing energy from the string.

3.5 Conservative vs Non-Conservative Forces

Let's begin by considering a ball near the surface of the earth. The ball begins three meters above the ground, falls, and bounces back up to its original height. How much work is done by gravity in this case? Well, work is always defined as the force of interest times the displacement times the cosine of the angle between the force and the displacement. In this case the force we're interested in is gravity, mg , so the work done is $mgd\cos(\theta)$. On the way down, the force is parallel to the displacement. The ball is moving down and the force is down, so the angle between the force and the displacement is zero degrees. The cosine of zero degrees is 1, so the work done by gravity as the ball travels down is positive mgd on the way up. However, the force is down and the displacement is up, so the force and the displacement are antiparallel, or the angle between them is 180 degrees. Now the cosine of 180 degrees is -1, and so the work done by gravity as the ball travels back up is $-mgd$. The total work done by gravity on the entire loop is the sum of the work done by gravity on the way down plus the work done by gravity on the way back up, which in this case is equal to zero.

The ball in this example is moving on what is called a closed path. The ball starts and stops in the same place. We've just seen that for this closed path, the work done by the gravitational force is equal to zero. It turns out that the work done by gravity on any closed path is equal to zero. This leads to a question: is this statement true for all forces? Is the work done by any force on any closed path always equal to zero?

To answer this question, let's think about a different force: friction. In this example, we have a 2kg box as it's dragged three meters across a table, with a coefficient of kinetic friction of 0.2, and back, and we're interested in the work done by the force of kinetic friction over this closed path. Once again, the work as always is the force times the displacement times the cosine of the angle in between. In this case, the force of interest is the force of kinetic friction, which we know to write as the coefficient of kinetic friction, μk , times the normal force. The box is not moving in the vertical direction, which I've called y in this example, so we know that the acceleration in the y direction is equal to zero. Thus, by Newton's second law, we know that the force in the y direction is equal to zero, the net force is equal to zero.

Thus, we can conclude that the normal force in this problem, is equal to the weight of the box, mg . Thus, we have the force of kinetic friction, $\mu k * mg$. As the box is dragged to the left, the force of friction is opposite the displacement, thus the angle between them is 180 degrees. The box is moving to the left, but the force is opposing the motion, the angle between them is 180 degrees. The cosine of 180 degrees is, once again, -1. And so, the work done by the force of friction as it moves to the left is $-\mu k(mgd)$. On the way to the right, the force is still opposite the displacement. Now the block is going to the right, but the friction is still opposing the motion, so the angle between the force and the displacement is still 180 degrees, which means that the work done by friction as the block goes to the right is also $-\mu k(mgd)$. Thus, we see in this case that the work done by the force of friction around this closed path, the box starts and stops in the same place, so this is a closed path, the work done on this closed path is $-2\mu k(mgd)$, which is not equal to 0.

Thus, it seems that we have two different types of forces. We have forces for which the work done over a closed path is always equal to 0, an example of this is gravity, we call such forces conservative. We have another type of force for which the work done

over a closed path is not equal to 0. The frictional force that we just saw is an example of this. Forces for which the work done over a closed path is not equal to 0 are called non-conservative forces. The definition of a conservative force is a force for which the work done by the force over a closed path is equal to 0. So, this is the statement you can use to determine if a force is conservative or non-conservative. Why do we care about this distinction between conservative and non-conservative forces? Because only conservative forces are associated with the potential energy.

To explore this idea, let's consider a block sliding down a frictionless ramp and a block just falling to the ground. Both blocks travel the same vertical distance h . In both cases, the work done by gravity is equal to the change in gravitational potential energy. The work done by gravity is written mathematically as a $-\Delta U$, or $-\Delta mgh$. The reason we can write the work done by gravity in terms of a potential energy is because gravity is a conservative force. Friction is not a conservative force; thus, we cannot write the work done by friction as the change in some type of frictional potential energy. So, let's summarize. We've seen that there are two classes of forces. Forces over which the work done over a closed path is zero, such as gravity, such forces are known as conservative, and for these types of forces we can describe the work done by the force as a change in the potential energy. $W = -\Delta U$. We also have forces for which the work done over a closed path is not equal to zero, such as friction. For these non-conservative forces, we cannot describe the work as in terms of a change in potential energy. There is no such thing as a potential energy for friction.

3.6 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 $K = \frac{1}{2}mv^2$, gravitational potential energy $U_g = mgy$, and the potential energy of a spring $U_s = \frac{1}{2}kx^2$. These different types of energy can be organized as in the chart in the **figure**. Collectively these types of energy are called *mechanical energy*.

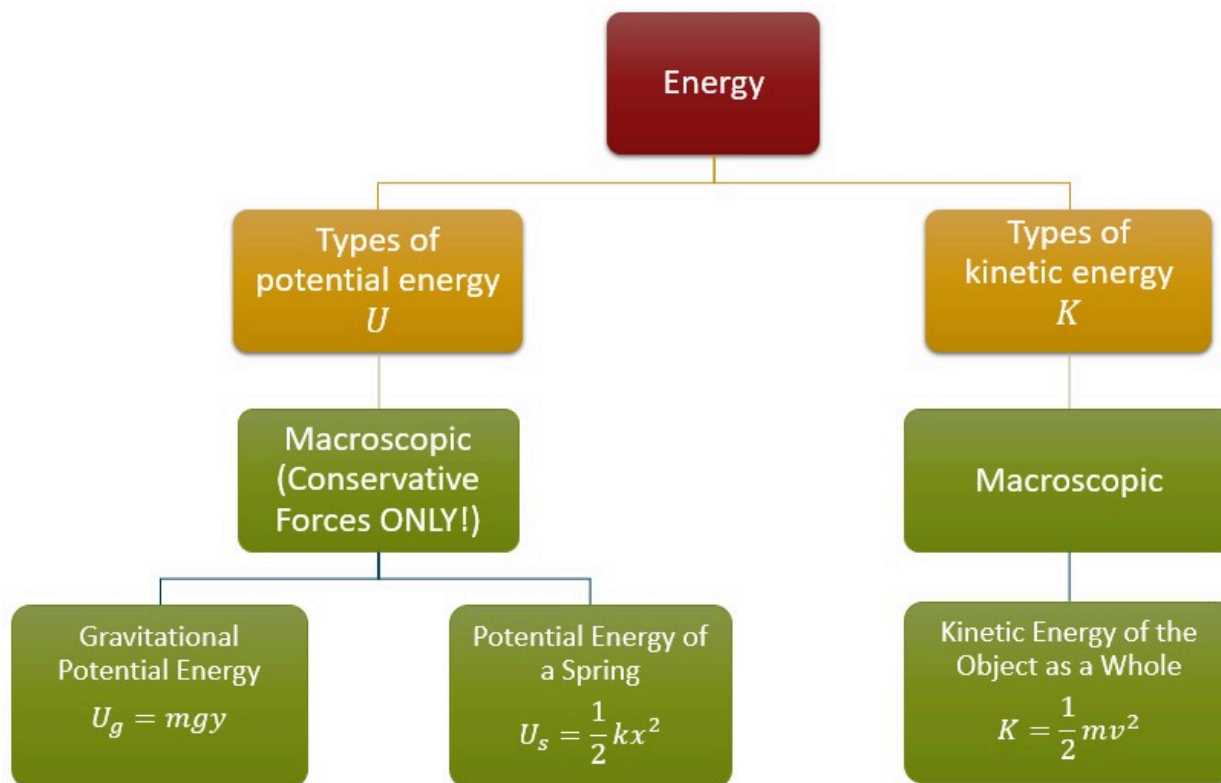


Figure 3.9 The different types of macroscopic (also called mechanical) energy

Glossary

gravitational potential energy: the energy an object has due to its position in a gravitational field

4 ENERGY OF CONSTITUENT ATOMS (MICROSCOPIC SCALE)

4.1 Introduction

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 $\Delta E = Q + W$ where the total energy E is still the sum of the potential energies U and kinetic energies K : $E = U + K$. 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 Q 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 U_{chem} . 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: K_{therm} is related to T . The relationships between these different forms of energy are shown in Figure 1. The total energy E at the microscopic scale is sometimes called the *internal energy* of the system.

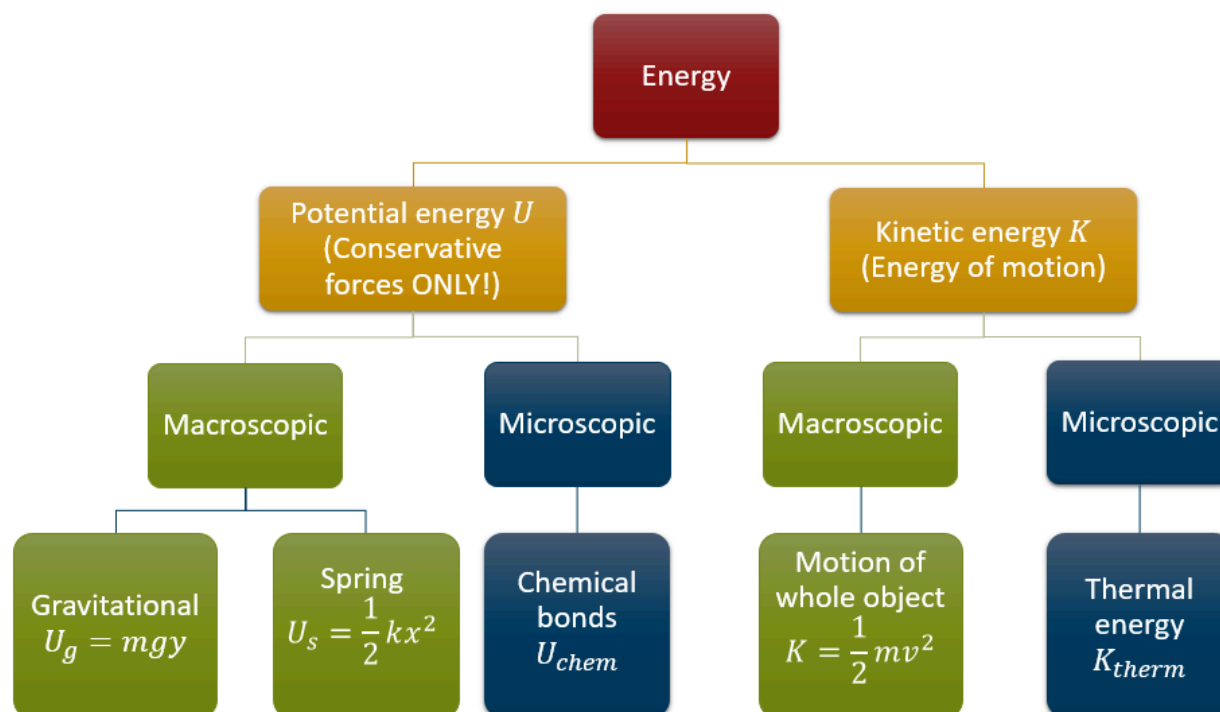


Figure 4.1 The different types of energy classified as microscopic and macroscopic

4.2 The 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 U_{chem} . 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 $\Delta H = 242 \text{ kJ/mol}$. 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 4.2 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!

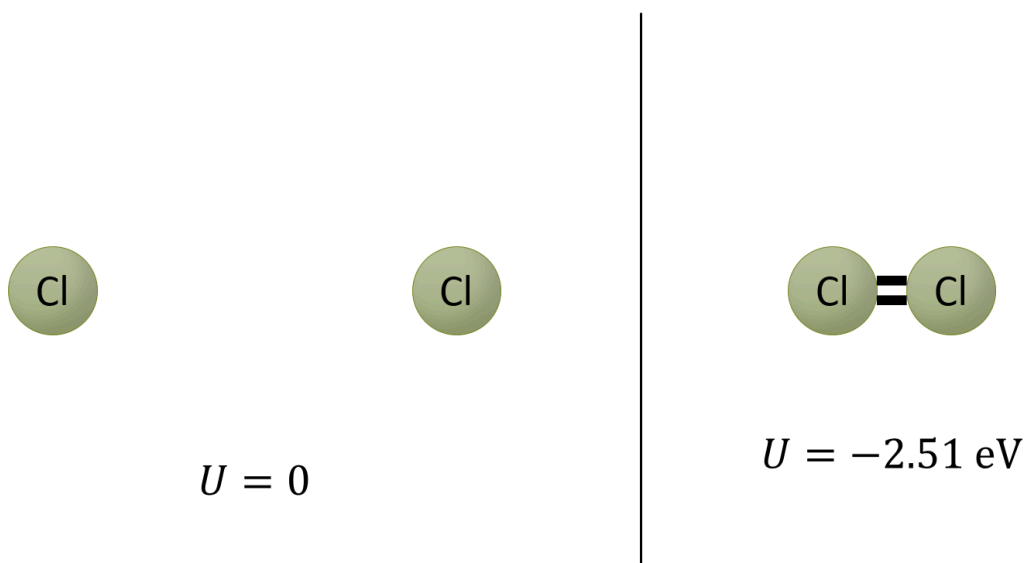


Figure 4.3 Two Cl atoms separated by a great distance have zero potential energy while two bonded Cl atoms have a potential energy of -2.51 eV. Remember, potential energy is due to the relative position of two objects, so it does not make sense to ask which atom in the bonded pair has the potential energy. 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.

4.3 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.

Example 4.1 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 monatomic H into 6.000g of H₂?

Solution

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

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

$$(6.000\text{gH}) \cdot \left(\frac{1\text{molH}}{1.008\text{gH}}\right) \cdot \left(\frac{6.022 \times 10^{23} \text{Hatoms}}{1\text{molH}}\right) = 3.584 \times 10^{24} \text{Hatoms}$$

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

$$\frac{3.584 \times 10^{24}}{2} = 1.79 \times 10^{24} \text{H}_2 \text{molecules}$$

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

$$(1.79 \times 10^{24} \text{H} - \text{Hbonds}) \cdot \left(\frac{4.52\text{eV}}{\text{H} - \text{Hbond}}\right) = 8.10 \times 10^{24} \text{eV}$$

which we can convert to Joules

$$(8.10 \times 10^{24} \text{eV}) \cdot \left(\frac{1.602 \times 10^{-19} \text{J}}{1\text{eV}}\right) = 1.30 \times 10^6 \text{J} = 1.30 \text{MJ}$$

Example 4.2 Bond energies expressed as enthalpies

The enthalpy of dissociation of $\text{CO}_2 \rightarrow \text{CO} + \text{O}$ is $\Delta H = +532 \text{ kJ/mol}$. How much *energy* is required to break up 20g of CO_2 into CO and O assuming the reaction occurs at constant pressure?

Solution

Also at https://youtu.be/SE_SN3plzbY (https://youtu.be/SE_SN3plzbY)

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

$$H = E + PV$$

At constant pressure, the change in enthalpy is

$$\Delta H = \Delta E + P\Delta V$$

Using the first Law of Thermodynamics $\Delta E = Q + W$

$$\Delta H = (Q + W) + P\Delta V$$

Recalling from section 13.8 the argument that, under constant pressure, the work W and $P\Delta V$ are of equal magnitude, but opposite sign, we have

$$\Delta H = Q$$

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

Now, we know that the *enthalpy* change $\Delta H = +532 \text{ kJ/mol}$ is the same as the *heat* $Q = +532 \text{ kJ/mol}$. Since the heat is positive we must add heat to make the process go; one mol of CO_2 going to CO and O will need 532 kJ of energy.

UMASS AMHERST Instructor's Notes

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 CO_2 in the problem. From the periodic table we see that the molar mass of CO_2 is:

$$m_{\text{C}} = 12.01 \text{ g/mol}$$

$$m_{\text{O}} = 16.00 \text{ g/mol}$$

$$m_{\text{CO}_2} = (12.01 \text{ g/mol}) + 2(16.00 \text{ g/mol})$$

$$m_{\text{CO}_2} = 44.01 \text{ g/mol}$$

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

$$(20 \text{ g CO}_2) \cdot \left(\frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \right) \cdot \left(\frac{532 \text{ kJ heat required}}{\text{mol CO}_2} \right) = 241 \text{ kJ}$$

We must add 241 kJ to break up 20g CO_2 into CO and O.

4.4 Kinetic energy at the Microscopic Scale

UMASS AMHERST Instructor's Notes

Your Quiz will Cover

- Using the relationship between kinetic energy, temperature, and degrees of freedom.

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.

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**.

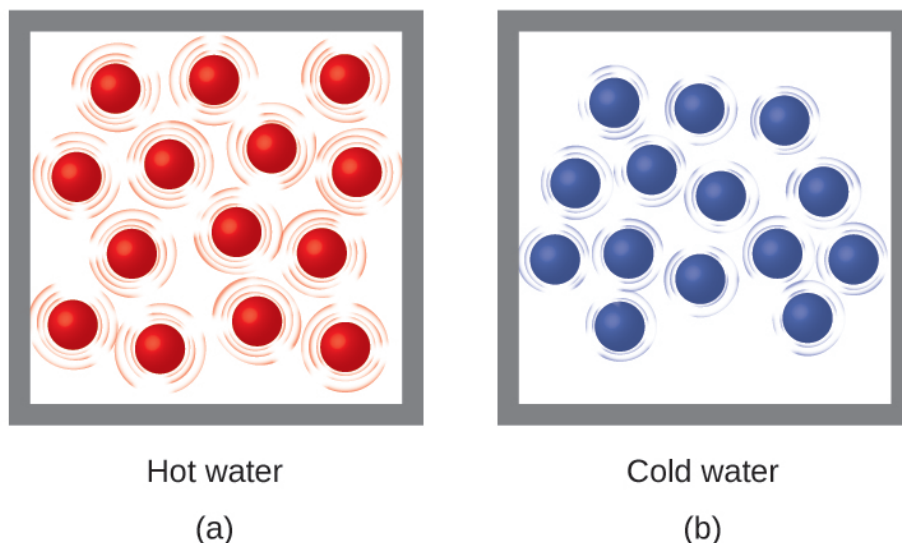


Figure 4.4 (a) The molecules in a sample of hot water move more rapidly than (b) those in a sample of cold water

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

$$k_B T = \frac{2\langle E \rangle}{n_{d.o.f.}}$$

or more commonly written

$$\langle K \rangle = \frac{n_{d.o.f.}}{2} k_B T.$$

In this expression $\langle K \rangle$ is the average energy, $n_{d.o.f.}$ is the number of degrees of freedom (places to put energy), and T is the temperature in *Kelvin*. The factor of 2 comes out of a formal calculus-based derivation, but as far as we are concerned, it is just there for convenience purposes. The symbol k_B is the Boltzmann constant introduced in 12.4 $k_B = 1.38 \times 10^{-23} \text{ J/K}$ 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.

Also at: <https://youtu.be/QXjZpUDIC6A> (<https://youtu.be/QXjZpUDIC6A>)

For an ideal gas, there were in fact three degrees of freedom, as I could either put energy into motion in the X direction, motion in the Y direction, or motion in the Z direction. So, three places I can put energy, three degrees of freedom. What about in a solid? In a solid, the atoms are not free to move around. Where can I put the energy in this case? I can clearly add energy to a solid, I can put it over a flame, what are the different microscopic places, different degrees of freedom that are possible in a solid?

We are going to consider the Einsteinian solid, where we treat each atom as connected to its neighbors in three dimensions by springs. In an Einsteinian solid, the atoms can't move anywhere, but they can vibrate on these springs that we're using to represent the atomic bonds. You might think, there are three bonds so there are three places I can put energy, I could put it in

any of the three bonds, and this is a very reasonable assumption. but not quite right. In fact, there are six degrees of freedom in a solid. There are two for each bond, not just one. When I add energy to a vibrating bond in a solid, there are two places I can put it. I can make the vibrations larger, increase their amplitude, or I can make it vibrate more quickly. Thus, there are two degrees of freedom per bond in a three-dimensional solid, and therefore, in a 3D solid, there are two degrees of freedom per bond, size of vibration and speed of vibration, times three bonds, gives us a total of six degrees of freedom for a three-dimensional Einsteinian solid.

In summary, on degrees of freedom, what you're looking at are the number of places within the individual atoms you can put energy. For an ideal gas, there are three places, I can put energy either into the motion in the X Direction, motion in the Y direction, or I could put energy into motion in the Z direction, and these three directions, as we've been discussing since the beginning of class are independent. For an atom in a solid, however, there are actually two degrees of freedom per bond, one for the size of the vibration and one for the speed of the vibrations. For a standard solid in three dimensions where each atom has three independent bonds, one in X, one in Y, and one in Z, there are 3 times 2, or 6, degrees of freedom.

Example using temperature, kinetic energy, and degrees of freedom

(a) What is the average kinetic energy of an ideal gas molecule at 20.0°C (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

$$\langle K \rangle = \frac{n_{dof}}{2} k_B T$$

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

$$\langle K \rangle = \frac{3}{2} k_B T$$

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

$$20^\circ \text{C} = 293.15 \text{K}$$

Now substituting it all together we have

$$\langle K \rangle = \frac{3}{2} (1.38 \times 10^{-23} \text{ J/K}) (293.15 \text{K})$$

$$\langle K \rangle = 6.07 \times 10^{-21} \text{ J} = 0.038 \text{ eV} = 38 \text{ meV}$$

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

$$\langle K \rangle = \left\langle \frac{1}{2} m v^2 \right\rangle.$$

Both $\frac{1}{2}$ and m are constants and can come out of the average

$$\langle K \rangle = \frac{1}{2} m \langle v^2 \rangle$$

$$\langle v^2 \rangle = \frac{2 \langle K \rangle}{m}$$

$$\langle v \rangle = \sqrt{\frac{2 \langle K \rangle}{m}}$$

The mass of helium is, from our periodic table, 4.002 amu which is $6.626 \times 10^{-27} \text{ kg}$. Putting this value and our result from (a) in our expression we get

$$v > \sqrt{\frac{2(6.07 \times 10^{-21} \text{ J})}{(6.626 \times 10^{-27} \text{ kg})}} = 1350 \text{ m/s}$$

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!

4.5 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 k_B

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

$$\langle K \rangle = \frac{n_{dof}}{2} k_B T,$$

then to get the total kinetic energy due to temperature K_{therm} , we take the average and multiply by the number of atoms N

$$K_{therm} = N \langle K \rangle.$$

Making a substitution we have

$$K_{therm} = N \left(\frac{n_{dof}}{2} k_B T \right).$$

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

$$N k_B = n R$$

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

$$K_{therm} = \frac{n_{dof}}{2} n R T$$

Example 4.3 Comparing 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 20°C. 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

$$U_g = mgh$$

$$U_g = (70\text{kg})(9.8\text{N/kg})(3\text{m})$$

$$U_g = 2000\text{J} = 2\text{kJ}$$

The thermal energy is given by

$$K_{therm} = \frac{n_{dof}}{2} n R T$$

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

$$(70\text{kg Pb}) \cdot \left(\frac{1000\text{g}}{\text{kg}} \right) \cdot \left(\frac{1\text{mol Pb}}{207.2\text{g Pb}} \right) = 337.8\text{mol Pb}$$

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

$$K_{therm} = \frac{6}{2} (337.8\text{mol}) \cdot \left(8.314 \frac{\text{J}}{\text{mol} \cdot \text{K}} \right) (293.15\text{K})$$

$$K_{therm} = 2.47 \times 10^6 \text{J} = 2.47\text{MJ}$$

Giving us a ratio of

$$\frac{U_g}{K_{therm}} = 8 \times 10^{-4}$$

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.

Example 4.4 Argon Gas

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 10°C assuming no work is done?

Solution

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

$$\Delta E = Q + W$$

which, given that no work is done goes to

$$\Delta E = Q$$

$$E_f - E_i = Q.$$

All of the energy in this case is thermal kinetic energy. Replacing E with K_{therm} , therefore, we have

$$K_{therm_f} - K_{therm_i} = Q.$$

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

$$\frac{n_{dof}}{2}RnT_f - \frac{n_{dof}}{2}RnT_i = Q$$

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

$$\frac{n_{dof}}{2}Rn(T_f - T_i) = Q.$$

Argon is an ideal gas, so $n_{dof} = 3$. We know the number of moles $n = 0.25mol$ and R is the gas constant

$R = 8.314 \frac{J}{mol \cdot K}$. Moreover, we know the temperature difference $T_f - T_i = 10^\circ C$. We, as always, need to convert to

Kelvins. However, the size of $1^\circ C$ is the same as $1K$, only the zero points are different. Therefore $T_f - T_i = 10K$.

Substituting these values in, we get

$$\frac{3}{2} \left(8.314 \frac{J}{mol \cdot K} \right) (0.25mol)(10K) = +31J$$

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

4.6 Connections to Chemistry

UMASS AMHERST Instructor's Notes

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, such as [here \(https://legacy.cnx.org/content/m51034/latest/\)](https://legacy.cnx.org/content/m51034/latest/), to be solved using the expression

$$mc\Delta T = Q.$$

In this expression, m is the mass of substance, ΔT is the change in temperature, and c is a quantity called *specific heat* with units $\frac{J}{kg \cdot K}$. If you have done these types of problems, then you have probably either solve for the specific heat c or looked it up in a table like the one below for the substance in which you were interested.

Table 4.1 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)

Substance	State	Specific Heat at constant volume (J/gK)
Helium	He(g)	3.12
Water	H ₂ O(l)	4.19
Ethanol	C ₂ H ₆ O(l)	2.3
Nitrogen	N ₂ (g)	0.743
Aluminum	Al(s)	0.87
Argon	Ar(g)	0.315
Iron	Fe(s)	0.46
Copper	Cu(s)	0.39

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

$$mc\Delta T = Q$$

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

$$\frac{n_{dof}}{2} R n (T_f - T_i) = Q$$

which after some rearranging becomes

$$n \left(\frac{n_{dof}}{2} R \right) \Delta T = Q.$$

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, M/M where M is the molar mass of the substance in g/mol (I know this seems dumb, but bear with me for just a few lines):

$$\left(\frac{M}{M} \right) n \left(\frac{n_{dof}}{2} R \right) \Delta T = Q \left(\frac{M}{M} \right)$$

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

$$(Mn) \left(\frac{n_{dof}}{2M} R \right) \Delta T = Q.$$

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

$$m \left(\frac{n_{dof}}{2M} R \right) \Delta T = Q$$

which when we compare with the calorimetry formula from chemistry

$$mc\Delta T = Q$$

we see that the only way this can work is if

$$c = \frac{n_{dof}}{2M} R$$

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 $n_{dof} = 3$. The molar mass for argon is 39.948 g/mol. Substituting into our expression we get

$$c = \frac{3}{2(39.948 \text{ g/mol})} \left(8.314 \frac{\text{J}}{\text{mol} \cdot \text{K}} \right) = 0.312 \frac{\text{J}}{\text{g} \cdot \text{K}}$$

which matches the table value exactly!

Solution for copper

Copper is a solid with $n_{dof} = 6$ and a molar mass of 63.546 g/mol. Thus, we expect the specific heat of copper to be

$$c = \frac{6}{2(63.546 \text{ g/mol})} \left(8.314 \frac{\text{J}}{\text{mol} \cdot \text{K}} \right) = 0.392 \frac{\text{J}}{\text{g} \cdot \text{K}}$$

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.

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