Chapter 1

Matter and Energy

The universe is made of matter and energy. Current models posit that the universe is approximately 68% dark energy, 27% dark matter, and 5% ordinary matter. Everything you can see and touch is part of the small part of the universe made of ordinary matter. Most science deals with ordinary matter and its interactions; highly trained theoretical physicists are currently debating the nature and effects of dark matter and dark energy.

What is this ordinary matter made of? All things (including the air around you) are made of atoms. Atoms are incredibly tiny — there are more atoms in a drop of water than there are drops of water in all the oceans.

Every atom has a nucleus that contains protons and neutrons. Orbiting around the nucleus is a cloud of electrons. However, the mass of the atom comes mainly from the protons and neutrons, since they are about 2000 times as massive as an electron! These three particles, protons, neutrons, and electrons, are called *subatomic particles*. (See figure 1.1.)

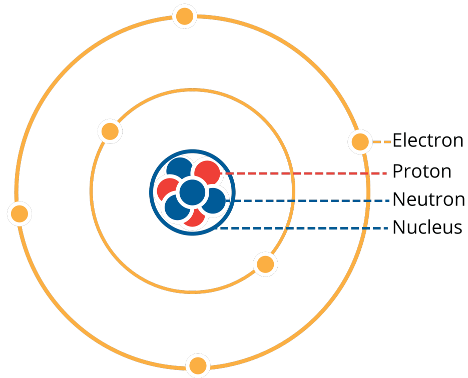


Figure 1.1

1.1 Atoms and Their Models

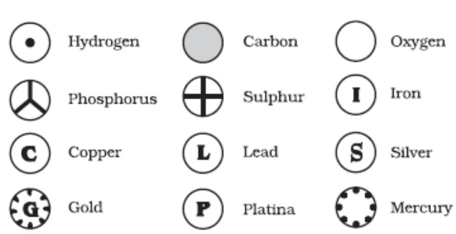


Figure 1.2

Over the history of science, there have been many ideas about the structure of atoms. This history is a good example of how science develops: unexpected results drive scientists to update their models, moving us closer and closer to a true model of the atom.

During his investigations into the behavior of gases, John Dalton noted that different elements combine in strict ratios. For example, he noted that nitrogen and oxygen combine in a 1:1 and 1:2 fashion, but no ratio in between.

This first model of the atom is rudimentary; each element is a unique atom, and those atoms cannot be subdivided. The atom is modeled as one large, solid, uniform, and neutral object. British physicist J.J. Thomson discovered that atoms could be split into a light, negatively charged particle and a heavier, positively charged particle (we now know this is the nucleus, the dense grouping of protons and neutrons in the center of an atom).

Suddenly, the atom went from neutral and indivisible to made of different types of charged particles. Further experiments by Ernest Rutherford showed the atom to be mainly empty space, further updating scientists' model of the atom. Subsequently, Bohr explained the phenomena of spectral lines (we will discuss this further in Sequence 2) by modeling electrons as being restricted to orbiting specific distances from the nucleus.

Figure 1.3

This is likely the model you are most familiar with seeing, and it is the one we will use most often in this text. The first figure shown in this chapter is a Bohr model: it shows the protons and neutrons in the nucleus, and models the electrons as moving in discrete orbits around the nucleus.

However, the Bohr model is slightly inaccurate. While it is a convenient model for thinking about atoms, in reality, electrons don't neatly orbit the nucleus. Scientists don't know exactly where an electron will be in relation to the nucleus, but they do know where it is most likely to be. They use a cloud that is thicker in the center but fades out at the edges to represent an electron's position (see figure 1.3).

While the cloud model is more accurate, we will use the Bohr model as it allows the viewer to easily and quickly assess the number and arrangement of electrons.

1.1.1 Classifying Atoms

We classify atoms by the numbers of protons they have. An atom with one proton is a hydrogen atom, an atom with two protons is a helium atom, and so forth (refer to the periodic table on page 44). We say that hydrogen and helium are *elements* because the classification of elements is based on the proton number. And we give each element an atomic symbol. Hydrogen gets H, helium gets He, oxygen gets O, carbon gets C, and so on. You can see an element's symbol on the periodic table.

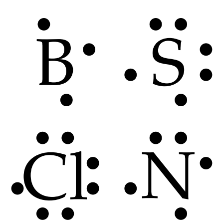


Figure 1.4

1.1.2 When Atoms Combine

When atoms of different elements combine, they make *compounds*. Compounds are substances made up of more than one element. Compounds can be *molecules* or *crystal lattices*. In the next section you'll learn *why* these different structures form.

There are many kinds of compounds. You know a few:

- Table salt is crystals made of Na^+ and Cl^- ions: a sodium atom that as lost an electron and a chlorine atom that gained an electron
- Baking soda, or sodium bicarbonate, is $NaHCO_3$.
- O_2 is the oxygen molecules that you breathe out of the air (air, a blend of gases, is mostly N_2.).
- Common quartz is SiO_2: silicon dioxide

The subscripts indicate what ratio of the elements are present in the compound. Each number indicates the ratio for the preceding element. A drop of water, H_2O, has twice as many hydrogen atoms as oxygen atoms.

Example: What is the ratio of elements present in Epsom salt?

Solution: Epsom salt, chemical name magnesium sulfate, has the chemical formula $MgSO_4$. Therefore, the ratio of Mg:S:O is 1:1:4.

Exercise 1 Ratios of Atoms in Molecules

Give the elemental ratio for each compound.

1. methane, CH_4
2. copper (II) sulfate, $CuSO_4$
3. glucose, $C_6H_{12}O_6$

Working Space

Answer on Page 11

1.2 Types of Matter

One way to classify matter is by the types of chemical bonds that hold a material's atoms together. The nature of these bonds, in turn, affects the properties of the material. For now, all you need to know is there are three types of chemical bonds: metallic, covalent, and ionic. Materials held together with the same type of bonds tend to have similar properties. For example, copper, bronze, iron, and steel (all containing metallic bonds) are all shiny, ductile, malleable, and good conductors of heat and electricity. On the other hand, Epsom salt and table salt form large crystals, have very high melting points, and are poor conductors of electricity in their pure form. These two substances (Epsom and table salt) both contain ionic bonds.

1.2.1 Covalent Compounds

Water is an example of a covalent compound: it is made of two hydrogen atoms covalently bonded one oxygen atom (see figure 1.5). The result is a water molecule. The different atoms cluster together because they share electrons in their clouds. This is the nature of a *covalent bond*: it is formed when atoms share electrons. Sometimes, electrons are shared evenly, but in water, they are shared unevenly. Oxygen is better at attracting electrons to itself than hydrogen, and so the shared electrons so they spend more time on the oxygen atom than the hydrogen atoms. As a result, the oxygen side of a water molecule has a slight negative charge, while the hydrogen atoms have a slight positive charge. These slight charges are represented with a lower case Greek letter delta, δ. When electrons are shared unevenly, we call this a *polar* covalent bond, because there are positive and negative poles at either end of the bond.

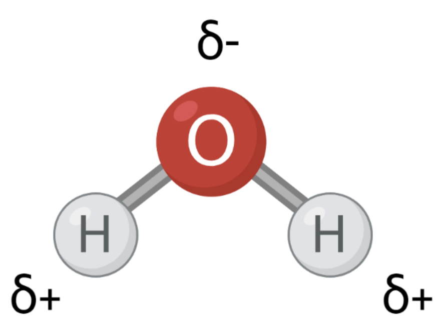


Figure 1.5

Whether electrons are shared evenly or unevenly is based on the elements' relative *electronegativities*. Electronegativity is simply a measure of how strongly an atom can attract electrons to itself. In general, elements on the right side of the periodic table are more electronegative than elements on the left side. When covalent bonds form between two elements of similar electronegativities, the electrons are shared evenly. We call this a *non-polar* covalent bond. Oil is an example of a non-polar covalent substance. Different oils have different ratios and structures, but all oils are made mainly of carbon and hydrogen, which have similar electronegativities. What happens when you try to mix oil and water? They don't mix well! This is due to the difference between their bond types. Polar substances, like water, mix best with other polar substances, while non-polar substances, like oils, mix best with non-polar substances. You'll learn more about why this is in Sequence 2.

For both polar and non-polar covalent bonds, the electrons are held tightly to the nuclei, even if they are shared among atoms. Those electrons don't move to another molecule: they move around within the molecule they are already a part of. Since electrons don't flow freely in covalent substances, they are also poor conductors of electricity. Covalent compounds also tend to have lower melting and boiling points compared to ionic compounds.

1.2.2 Ionic Compounds

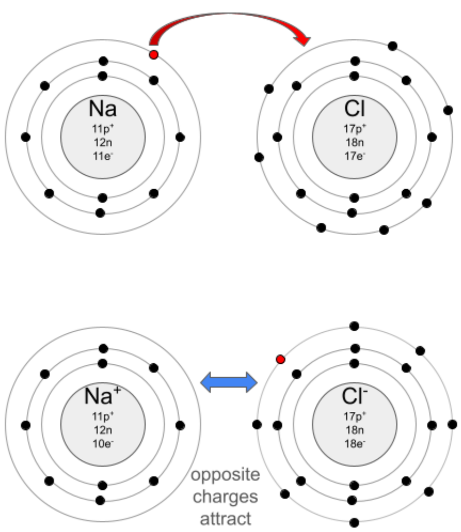
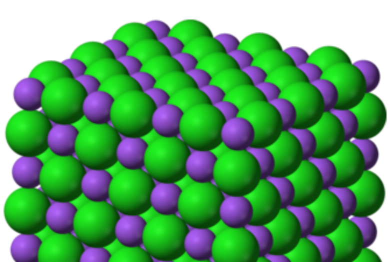


Figure 1.6

Ionic bonds are the electrical attraction between opposite-charged ions. When a neutral atom gains or loses an electron it becomes an *ion* (a charged atom), and oppositely-charged ions are attracted to each other. Which atom gets the electron and which loses it is based on their electronegativities: the more electronegative atom steals one or more electrons from the less electronegative atom. There are also polyatomic ions: groups of atoms held together with covalent bonds that have an overall charge (figure ... shows the Lewis dot structure of a phosphate polyatomic ion, as an example). For now, we'll focus just on ionic bonds between monoatomic ions, like in table salt. You'll learn more about polyatomic ions and the compounds they form in Sequence 2.

Let's examine how a simple ionic compound is formed: sodium chloride, also known as table salt, is made up of sodium and chlorine atoms (see figure 1.6). When sodium and chlorine come in contact with each other, electrons move from the sodium to the chlorine, making a sodium *cation* (positively-charged ion) and a chloride *anion* (negatively-charged ion). Yes — chlor*ide* is correct! When naming an anion, the ending of the element name changes to *-ide*. Once the sodium cation and chloride anion are formed, their opposite charges attract them to each other, like north and south magnet poles.



When there are many, many sodium and chloride ions around, they spontaneously arrange themselves in a pattern, giving ionic compounds their characteristic crystal structure (see figure 1.7). Because the electrons are tightly held be each ion, ionic substances don't conduct electricity well as solids. The atomic crystal lattice also determines the shape of the macroscopic crystals. Salt crystals are generally cubic, while other crystals (like quartz) form hexagonal prisms. You'll learn how to predict the atomic and macroscopic crystal structure of different compounds in Sequence 2.

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Modeling Ionic Compounds

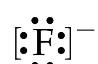
You can represent ions with Lewis dot diagrams by adding or subtracting electrons from the model. Since electrons are negative, anions have gained electrons while cations have lost them. Additionally, the ion is in brackets and the overall charge is indicated outside the top right corner of the brackets.

Example: Create Lewis dot diagrams for Na^+, F^-, O^{2+}, and Mg^-.

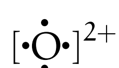
Solution: Sodium is in column 1, so a neutral sodium atom has 1 valence electron. A +1 charge means it has lost 1 electron, leaving zero.

$$[\mathrm{Na}]^+$$

Fluorine is in column 17 and has 7 valence electrons when neutral. The anion F^- has gained one electron, for a total of 8.



Neutral oxygen has 6 valence electrons, therefore O^{2+} has 4.



Neutral magnesium has 2 valence electrons, so Mg^- has 3 valence electrons.

$$[\cdot\dot{\mathrm{M}}\mathrm{g}\cdot]^{-}$$

To make Lewis dot diagrams of ionic compounds, you show both ions in the ratio given in the formula. For example, a Lewis dot diagram of $MgCl_2$ would show one magnesium cation and two chloride anions.

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1.2.3 Metallic Compounds

You may already know that metals (both pure and alloyed) are excellent conductors of electricity and heat. This is a consequence of their metallic bonds. In pure metals and alloys, the outermost layer of electrons can move freely from one atom to the next. As a result, at the atomic level, metals are best characterized as a lattice of cations surrounded by a "sea of electrons" (see figure fixme metallic bonding figure)

The free-flowing sea of electrons in pure and alloyed metal means the cation lattice can be rearranged without breaking the metallic bond. As a result, metals are ductile (able to be drawn out into a wire) and malleable (able to be hammered into a new shape without cracking or breaking). (fixme figure showing deformation of cation lattice in bending metal)

Metals can be pure, like copper or iron, or *alloys*, like bronze or steel. *Alloys* are mixtures between two or more elements where at least one element is a metal. Steel is an alloy of iron and carbon; bronze is an alloy of copper and tin. Alloying a metal changes its properties because of the change in the cation lattice (see figure fixme pure vs alloyed lattice and structural changes).

1.3 Energy and Work

Energy is defined as the ability to do work, but what does this mean? First, we need to understand what *work* is. When you lift an object into the air, you are doing work on that object. When water turns a turbine in a hydroelectric plant, the water is doing work on the turbine. And when you hit the brakes on your car, the brake pads are doing work on the tires (albeit, negative work). *Energy* is being transferred between these pairs of objects when work is done.

Some everyday examples of energy include:

1. The Calories in your food
2. The light from the Sun

3. Heat in the Earth's mantle
4. The motion of a spinning wheel

Some types of energy are easy to visualize, while others are not. Energy is what moves from one object to another when work is being done. When you lift something, the energy stored in your body (in the form of sugar and fat) is transferred to the object, accelerating it upwards. Your body continues to transfer energy as you lift the object against gravity. When you've lifted it as high as you can, most of the energy your body lost (we call this "burning Calories" colloquially) is stored as *potential energy*, due to the object's height.

Another example is a simple circuit connecting a battery and a light bulb. The battery has stored potential energy. When the circuit is complete, the potential energy in the battery is transferred to electrons in the light bulb, causing them to move and gain kinetic energy. In the light bulb's filament (we are referencing old, non-LED light bulbs here!), the electrons encounter resistance, which slows them down. The energy the electrons lose in this process is being transformed into light and heat, lighting your room.

The Work-Energy theorem explains the relationship between work and energy, and we will introduce the theorem and use it to explain energy transfer in a subsequent chapter.

1.4 Mass-Energy Equivalence

You've probably seen the equation

$$E = mc^2$$

E is energy, m is mass, and c is the speed of light in a vacuum (about $3 \times 10^8 \frac{m}{s}$). So far, we've been discussing matter and energy separately. This equation shows that matter can be converted to energy, and vice-versa. This is the source of the light energy emitted by the sun.

The Sun is mostly hydrogen. At the very center, it is so hot and dense that the nuclei of hydrogen atoms are fused together to form helium atoms, a process called the proton-proton chain reaction. The actual reaction involves several steps and is more complicated, but the overall process can be summarized as:

$$4\text{H} \rightarrow \text{He}$$

If every hydrogen atom has a mass of approximately $1.6735575 \times 10^{-27}$ kg and every helium atom has a mass of approximately $6.6464731 \times 10^{-27}$ kg, how much energy is released when one atom of helium is created? First, notice that one helium has less mass than 4 hydrogen atoms:

$$4 \times \left(1.6735575 \times 1{-}^{27}\right) - 6.6464731 \times 10^{-27} = 4.77569 \times 10^{-29}$$

Now, we can use $E = mc^2$ to find out how much energy is equivalent to 4.77569×10^{-29} kg:

$$E = \left(4.77569 \times 10^{-29}\text{kg}\right)\left(2.99792458 \times 10^8 \frac{\text{m}}{\text{s}}\right)^2$$
$$E \approx 4.292 \times 10^{-12}\text{joules}$$

All of these numbers are very small and hard to visualize. We could ask this: if 1 kilogram of hydrogen (about enough to fill a standard beach ball) were fused to make helium, how much energy would be released? For every kilogram of hydrogen that enters the proton-proton chain reaction, about 0.02854 kilograms of mass are converted to energy (the mass of about 5 quarters).

$$E = (0.02854 \text{ kg})\left(2.99792458 \times 10^8 \frac{\text{m}}{\text{s}}\right)^2 \approx 2.5647 \times 10^{15} \text{ joules}$$

This is more than 700,000,000 kWh (kilowatt hours); the average US household uses only 30 kWh per day. Fusing one kilogram of hydrogen releases enough energy to power an average US home every day for *over 65,000 years*! This relatively huge release of energy is why scientists continue to research nuclear fusion energy sources. The nuclear power plants currently running around the world today rely on *nuclear fission*: the splitting apart of atoms, the opposite process of *nuclear fusion*. Nuclear fission releases much, much less energy per kilogram of input material than nuclear fusion, and thus stable, affordable nuclear fusion power plants remain a ”holy grail” of scientific research.

1.5 Conclusion

We have seen that the universe is made of dark energy, dark matter, and ordinary matter. Ordinary matter is made of atoms, which can be classified based on their number of protons. Atoms combine in different ways to make compounds, and the manner of combination (ionic, covalent, or metallic bonding) affects the macroscopic properties of the substance. Energy allows matter to do work, and work is the transfer of energy.

Matter and energy do share a fundamental property: they are both conserved. Neither matter nor energy can be created or destroyed. This means the total amount of ordinary matter and energy in the universe is constant. This great scientific truth *something about its application* In the next chapter,

Appendix A

Answers to Exercises

Answer to Exercise 1 (on page 3)

1. C:H = 1:4
2. Cu:S:O = 1:1:4
3. C:H:O = 6:12:6 = 1:2:1

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