

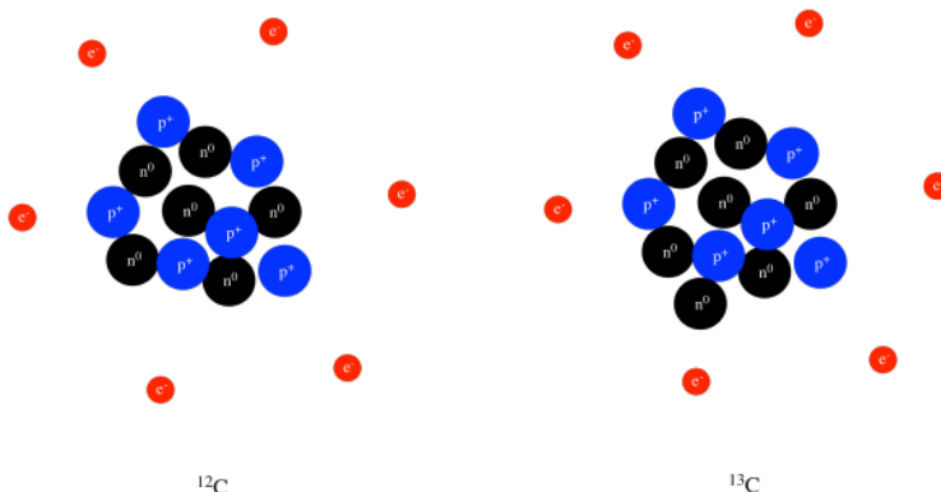
1.2: Old Quantum Mechanics- Basic Developments

Today, we know that atoms contain protons, neutrons and electrons. The protons have significant mass and a positive charge and are found in the nucleus of the atom. The neutrons have mass but no charge and are also found in the nucleus. The electrons have negative charge and very little mass and are found outside the atom's nucleus. The weight of an atom in atomic mass units is approximately the sum of its protons and neutrons, since the electrons don't have much mass.

For example, a typical carbon atom has six protons and six neutrons, and it has an atomic weight of 12 amu. A carbon also has six electrons, but they are so small that they don't contribute to carbon's weight.

- An element is defined by the number of protons in its nucleus.
- The number of protons in an atom is equal to the number of electrons, to balance the charge.

Some carbon atoms have an extra neutron or two, so carbon may have an atomic weight of 13 or even 14 amu. However, a carbon atom can't have an extra proton; an extra proton would make it a nitrogen atom. It is the six protons that make the atom behave like carbon. Many other elements may have slightly different atomic weights, depending on how many neutrons are found in the nucleus. Different atoms of the same element with different weights are called *isotopes*. For example, ^{12}C , ^{13}C and ^{14}C are all isotopes of carbon. They all have six protons, but different numbers of neutrons, as seen in a model of ^{12}C and ^{13}C , below.



- Neutrons are also in the nucleus.
- A neutron has a mass similar to a proton, but has no charge.
- Compared to protons and neutrons, the mass of an electron is very small.

? Exercise 1.2.1

An element's atomic number is just the number of protons in an atom of that element. Given the following atomic numbers and atomic weights, identify the number of protons, neutrons and electrons in an atom of the element.

- oxygen: atomic number = 8, atomic weight = 16
- phosphorus: atomic number = 15, atomic weight = 31
- zinc: atomic number = 30, atomic weight = 65
- gold: atomic number = 79, atomic weight = 197

Answer a:

O: 8 protons, 8 electrons, 8 neutrons

The atom is neutral overall, so the number of positively charged protons is equal to the number of negatively charged electrons. The atomic weight is provided almost entirely by the protons and neutrons, so the number of protons plus number of neutrons equals the atomic weight.

Answer b:

P: 15 protons, 15 electrons, 16 neutrons

Answer c:

Zn: 30 protons, 30 electrons, 35 neutrons

Answer d:

Au: 79 protons, 79 electrons, 118 neutrons

? Exercise 1.2.2

If a proton's mass is 1.67×10^{-27} kg and the mass of an electron is 9.11×10^{-31} kg, how many times heavier is a proton than an electron?

? Exercise 1.2.3

If carbon in nature is about 99% ^{12}C and 1% ^{13}C , then what is the average weight of a carbon atom?

Answer

Carbon: $[(99 \times 12 \text{ amu}) + (1 \times 13 \text{ amu})] / 100 = [1,188 + 13 \text{ amu}] / 100 = 1,201 \text{ amu} / 100 = 12.01 \text{ amu}$

? Exercise 1.2.4

Note that ^{14}C is even rarer than ^{13}C , because ^{14}C is converted into ^{14}N via radioactive decay. In that event, a high-energy electron is emitted from the ^{14}C nucleus. Explain how that emission must convert the carbon into a nitrogen, and indicate how many protons and neutrons are found in the resulting nucleus.

Answer

The negatively charged electron's departure from the nucleus leaves behind a positive charge. A neutron is converted into a proton. The overall atomic weight remains the same, but the atom ends up with one more proton and one more electron. A ^{14}C is converted into a ^{14}N .

? Exercise 1.2.5

Magnesium in nature is found in three major isotopes. It is nearly 79% ^{24}Mg , about 11% ^{25}Mg and 12% ^{26}Mg . What is the average weight of a magnesium atom?

? Exercise 1.2.6

Chlorine in nature is found in two major isotopes: ^{35}Cl and ^{37}Cl . If the average atomic weight of chlorine is about 35.5, what percentage of each isotope is found in nature?

Answer

Suppose y is the decimal fraction of ^{35}Cl and z is the decimal fraction of ^{37}Cl .

$$35.5 \text{ amu} = (y \times 35 \text{ amu} + z \times 37 \text{ amu})/100 \text{ but } y + z = 1$$

$$35.5 \text{ amu} = (y \times 35 \text{ amu} + (1-y) \times 37 \text{ amu})$$

$$35.5 \text{ amu} = y \times 35 \text{ amu} - y \times 37 \text{ amu} + 37 \text{ amu}$$

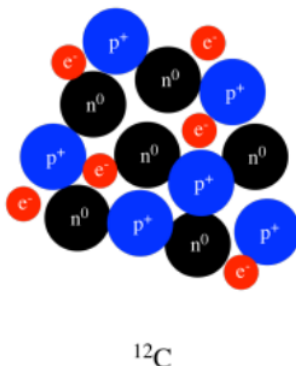
$$(37-35) \times y \text{ amu} = 37 \text{ amu} - 35.5 \text{ amu}$$

$$2 \times y \text{ amu} = 1.5 \text{ amu}$$

$$y = 0.75 \text{ (75\% } ^{35}\text{Cl)}$$

$$z = 0.25 \text{ (25\% } ^{37}\text{Cl)}$$

A number of developments at the beginning of the twentieth century led to our current understanding of the structure of atoms and molecules a hundred years later. At that time, some people thought protons, neutrons, and electrons were lumped together in the atom. This view of the atom was called the "pudding model" of the atom.



Ernest Rutherford first proposed that an atom contains a very small, positively charged nucleus surrounded by empty space. The electrons orbited far away from the nucleus.

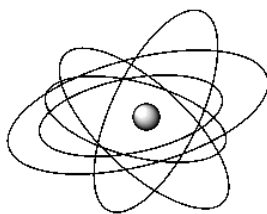


Figure 1.2.1: Rutherford's model of an atom: a nucleus with electrons far away, and lots of empty space.

Rutherford was explaining the result of an experiment in which alpha particles (positively charged helium ions) were fired at a gold foil. Most of these particles passed through the foil easily, suggesting there was a lot of empty space in the material. However, some of the particles bounced directly back, having collided with the small, highly charged nuclei. The positive alpha particles were powerfully repelled by the positive nuclei, because like charges repel each other. He didn't really know much about the location of the electrons, the negatively charged particles in the atom, but believed they orbited the nucleus like planets around the sun.

Why weren't the electrons found in the nucleus? If electrons are attracted to protons, it seems like that's where they should be. Niels Bohr suggested that electrons are found only in specific, allowed orbits at different distances from the nucleus.

That conceptual leap to specific, allowed orbits marks the introduction of quantum mechanics into the understanding of the atom. Quantum mechanics is based on the idea that on a very small scale, many properties only have specific values (like 1, 2, 3...) instead of any value at all (like all the possible fractions between these integers). In other words, in the world around us, we usually view things like walking up a ramp. We can heat a pot of water just a little bit warmer, and just a little bit warmer than that, and so on. On the atomic scale, however, the world is more like walking up a set of stairs. Maybe you could heat the water to 30 °C or 40 °C, but heating to 35 °C would be impossible, because heat only comes in 10 degree packages. That is, in fact, how the quantum world really works, but on the human scale, the steps involved are so tiny that we cannot notice them.

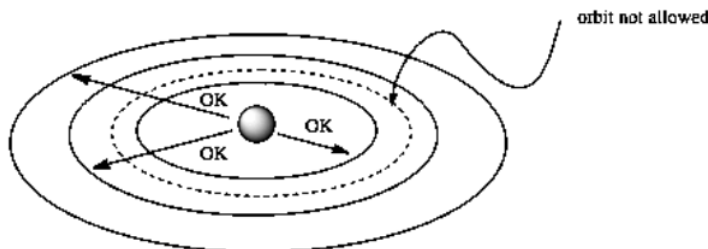
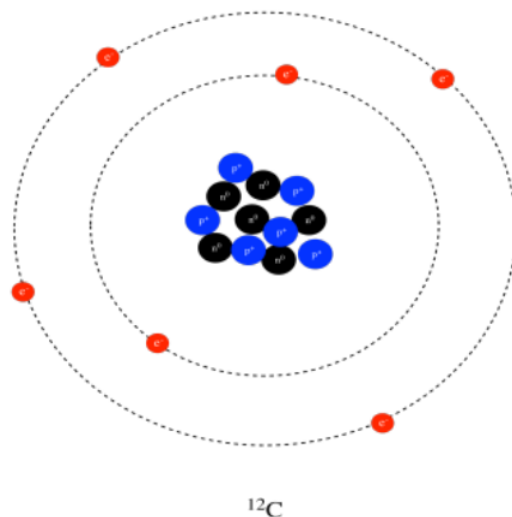
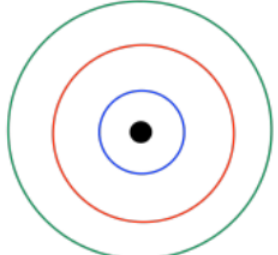


Figure 1.2.2: In the Bohr model of an atom, electrons could be found only at certain allowed distances from the nucleus.
Diagram showing that orbits must be at specific distances from the nucleus, not in-between two orbits.

Bohr's model was also consistent with the earlier idea of the periodic table of the elements. The idea is that electrons are found in different "shells" that are each further and further from the nucleus. Each of those shells corresponds roughly to a different row in the periodic table. Hydrogen and helium have electrons only in the first shell, and we see those two elements in the first row of the periodic table. Carbon and oxygen's outermost electrons are found in the second shell, so they show up in the second row of the periodic table.



Each row in the periodic table corresponds to an outer layer of electrons that are found further from the nucleus than the outermost electrons in the row before it. We are going to see eventually that there is a further variation on this idea, but it is still pretty much the way we see the periodic table today. Hydrogen starts the first shell, lithium the second, sodium (the Latin, natrium) the third, potassium (Latin, kalium) the fourth, and so on.



1	H			He
2	Li Be		B C N O F Ne	
3	Na			
4	K	3 Sc Ti V Cr Mn Fe Co Ni Cu Zn		
5	Rb			
6	Cs			
4				
5				

The variation we are going to see involves that dip in the middle of the periodic table. Scandium through zinc have outer electrons that are only in the third shell, not the fourth. The third and the fourth shell overlap a little bit, so that some electrons actually start to go into the fourth shell (as in potassium and calcium), then finish filling the third shell across the transition metals. The reasons for that also have to do with quantum mechanics, but we will need to learn a little more about energy and waves before we see why.

Bohr showed that electrons might be found in specific orbits around the nucleus. He also showed that electrons in these different orbits have specific amounts of energy. By doing this mathematically, he was offering an explanation to an important problem. People knew that atoms can absorb energy (they can be heated in a flame, for example) and give the energy back again in the form

of light. Rather than give off light of all colors when excited, atoms only give off very specific colors. For example, heating lithium salts in a flame produces a red color, but heating sodium salts produces an orange color, whereas potassium salts produce a purple color, and so on.

These colors can be separated and studied using a prism. When people did that, they found that a given atom does not produce just one pure color of light, but several different ones. When separated by a prism, the light given off by an excited compound could be seen against a dark surface as several different, colored lines. These were called emission lines.

It had been known since the early 1800's that light had wave properties, and that light of different colors had different wavelengths. For example, red light consists of electromagnetic waves, with a wavelength of about 700 nm, but blue light's wavelength is about 450 nm. That means a color can actually be measured numerically. Because of that fact, people can look for mathematical relationships between the emission lines observed for different atoms. Those mathematical relationships may reveal something about the atoms themselves.

Furthermore, it was known that different wavelengths of light corresponded to different amounts of energy. In one of the first developments in quantum mechanics, Max Planck in 1900 proposed that light travels in bundles called photons. Although they are particles, these photons do have wave properties. The amount of energy in a photon of light corresponds to its wavelength.

By proposing that electrons could be found only in specific orbits, specific distances away from the nucleus, Bohr was trying to explain observations from atomic spectroscopy reported by another scientist named Rydberg. Rydberg had found a mathematical relationship between the wavelengths of these emission lines. Bohr thought that, when energy was added, electrons could be excited from one energy level (or orbit) to a higher one. When the electron relaxed back to its original orbit, it gave off the energy it had gained in the form of light. The specific emission lines occur because electrons are found at very specific energy levels in an atom, so a drop from one level to another always produces the same amount of light energy. That specific amount of light energy has a specific color.



Figure 1.2.3: The correspondence between color, wavelength and energy.

Left: High energy blue light with short wavelength. Right: low energy red light with long wavelength.

Bohr then used the mathematical relationships describing electrostatic attraction and centripetal force to show that his model of the atom was consistent with Rydberg's relationship. In fact, he could use his model to predict the emission lines of an atom.

? Exercise 1.2.7

Bohr's explanation of atomic structure built on Rydberg's observation of a numerical series in spectral emission lines. Solving a series involves finding a pattern in numbers. Find the patterns among the following sequences of numbers, and predict the next number in the sequence.

- 1, 2, 3, 4...
- 2, 4, 6, 8...
- 3, 5, 7, 9...
- 1, 4, 9, 16...
- 2, 4, 8, 16...
- 1, $1/2$, $1/4$, $1/9$...

Answer a:

1, 2, 3, 4, 5... a series of whole numbers, or n .

Answer b:

2, 4, 6, 8, 10... a series of even numbers, or $2n$ (because every even number is two times another number).

Answer c:

3, 5, 7, 9, 11... a series of odd numbers, or $2n + 1$ (because every odd number is one more than some even number).

Answer d:

1, 4, 9, 16, 25... a series of squares, or n^2 .

Answer e:

2, 4, 8, 16, 32... a series in which each number is double the last number, or 2^n .

Answer f:

1, $1/2$, $1/4$, $1/9$, $1/16$... a series of reciprocals of squares, or $1/n^2$.

? Exercise 1.2.8

Bohr's idea depended partly on the use of Coulomb's Law of electrostatic attraction. Coulomb's law is expressed mathematically as follows:

$$F = \frac{kq_1 q_2}{r^2} \quad (1.2.1)$$

in which F is the attractive force between two charged particles, q_1 and q_2 are the charges on the two particles, r is the distance between the two particles and k is a constant. A large value of F means that the charges are strongly attracted to each other.

- Suppose q_1 is the charge on the nucleus of an atom and q_2 is the charge on an electron. What happens to the force of attraction between an electron and the nucleus when the charge in the nucleus increases? Explain.
- Suppose r is the distance from the electron to the nucleus. What happens to the force of attraction between an electron and nucleus when the electron gets further from the nucleus?
- Using the ideas of Coulomb's law, compare the attraction of an electron to the nucleus in a helium atom versus a hydrogen atom.

Answer a:

When q increases, F increases. The increase is linear: if q_1 doubles, F doubles. As the charge in the nucleus gets larger, the force of attraction gets larger.

Answer b:

When r increases, F decreases. The decrease is nonlinear: if r doubles, F drops by a factor of four, rather than a factor of two. As the distance from the nucleus gets longer, the attraction to the nucleus drops sharply.

Answer c:

Hydrogen and helium are both in the first row of the periodic table. To a rough approximation, the distance between nucleus and electron is similar in these two atoms. However, helium has a charge of $2+$ in its nucleus, compared to the $1+$ charge in the nucleus of a hydrogen atom. As a result, the attraction of an electron on helium to the nucleus would be about twice as great as the attraction of an electron on hydrogen to its nucleus.

Helium's electrons are much more tightly held than hydrogen's.

The situation is really much more complicated than that. For example, if helium's electrons are more tightly attracted to the nucleus than hydrogen's, then helium's electrons ought to be pulled closer to the nucleus than hydrogen's. That means helium's electrons are held even more tightly than we at first thought.

Another complicating factor is that helium has two electrons, whereas hydrogen has only one. An electron may be attracted to the nucleus, but electrons repel each other. That second electron in helium should offset the extra attraction to helium's more positive nucleus. That means helium's electrons may be less tightly attracted than we originally thought.

However, the effect of the second electron is much smaller than it first appears. That's because the second electron could be anywhere around the helium atom. It has a 50% chance of being further away from the first electron than that positively charged nucleus. The farther it is away, the smaller its influence. Helium's electrons are definitely more strongly attracted to the nucleus than are hydrogen's, but it is difficult to say exactly how much more without the help of some more sophisticated tools.

? Exercise 1.2.9

Max Planck described the energy of a photon using the following relationship:

$$E = h\nu \text{ or } E = \frac{hc}{\lambda}$$

In which E = energy; ν = frequency; λ = wavelength; c = speed of light; h = Planck's constant

- What happens to the energy of light as its wavelength gets longer?
- What happens to the energy of light as its frequency gets higher?

Answer a:

As wavelength gets longer (value of λ increases), energy decreases.

Answer b:

As frequency gets higher (value of ν increases), energy increases.

Other people were familiar with these ideas and already knew about the relationship between light and energy. Bohr's model of the atoms put all of these ideas together to successfully explain a specific atomic property:

color = wavelength = energy of light = energy between electron levels.

In other words, an excited electron can drop back to its original orbit by giving off a photon with an energy exactly the same as the difference in energy between the two orbits ("excited state" and "ground state" orbits).

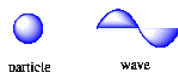


Figure 1.2.4: An electron can be thought of as both a particle and a wave.

However, Bohr did not explain why electrons would be found at specific energy levels in the first place. Louis de Broglie, a historian-turned-physicist, solved this problem with the idea of wave-particle duality. de Broglie put together the following ideas:

- All moving particles have wave properties.
- Electrons move around the nucleus and they have wavelengths.
- To maintain a complete standing wave along its orbit, an electron can only adopt orbits of specific circumferences. Otherwise, one end of the wave would not meet up with the other end, and it would interfere with itself.
- Orbits with specific circumferences have specific radii.
- Electrons are found at specific distances from the nucleus, but not at other distances.

One way to illustrate why an electron might have only certain allowed orbits is via the "particle in a box", a basic concept from quantum mechanics. If a particle has wave properties, then it has a wavelength. Its wavelength depends on certain conditions. By analogy, if you take a guitar string and attach it to the ends of a box, the string can only vibrate at certain frequencies. That's how guitarists can change the note played on a guitar string. By pressing one end of the string against a fret on the guitar neck, the length of the string is changed, and so is its allowed wavelength, so it makes a different sound.

The string can't move at the two points where it is held. That means the wave has to form in such a way that it returns to the same position at both ends. Because of that, certain wavelengths won't work, because the wave won't be able to return to that correct position at the far end.

Furthermore, the allowed wavelengths of a guitar string also depend on the thickness of the string. As a result, there are two conditions that control the tone that is played: which of the six guitar strings is plucked, and where the string is held against the frets.

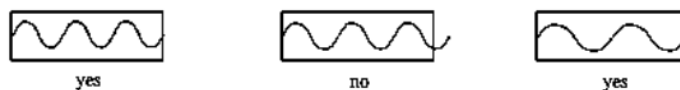


Figure 1.2.5: A particle in a box is allowed only certain wavelengths, based on the dimensions of the box.

Left: a wave in a box that does not extend outside the box, labelled "yes". Middle: a wave that extends outside the box, labelled "no". Right: a wave with a longer wavelength that does not extend outside the box, labelled "yes".

The same thing is true with very small particles that have wave properties. These particles can have only certain wavelengths that fit their surroundings. An electron has some property, analogous to the thickness of a guitar string, that limits its possible wavelengths. Given those limits, there are only certain orbits allowed the electron. If its orbit does not have the right circumference, the electron will not be able to form a complete wave along that orbit.

These ideas ushered in a revolution in science. Quantum mechanics is a very powerful tool. It can be used to accurately predict how molecules will behave. Unfortunately, the mathematics involved in quantum mechanics are one or two math courses beyond what most introductory chemistry students are familiar with. Even so, a qualitative feel for some of the consequences of quantum mechanics is important enough that we should explore it.

? Exercise 1.2.10

Complete the table with the appropriate information.

Table for Exercise 1.2.10

Element Symbol	Atomic Number	Mass Number	Number of Protons	Number of Neutrons	Number of Electrons	Charge
H		1			1	
H		1			0	
H		1			2	
H		2			1	
H		3			1	
	4	9			2	+2
	6	12		6	6	
	12			13	12	0
	43			55	43	
	20	40				+2
Si	14	28				0
			19	28		+4
Fe			26	30	23	
	35			44		-1
K			22	17	21	
	15			15		0
	13	27				+3
S				16	16	0
Pd		106	46			+1
	24			28	21	

Element Symbol	Atomic Number	Mass Number	Number of Protons	Number of Neutrons	Number of Electrons	Charge
			50	68	50	
Hg	80			120	79	
	79			118	78	

Answer

Element Symbol	Atomic Number	Mass Number	Number of Protons	Number of Neutrons	Number of Electrons	Charge
H	1	1	1	0	1	0
H	1	1	1	0	0	+1
H	1	1	1	0	2	-1
H	1	2	1	1	1	0
H	1	3	1	2	1	0
Be	4	9	4	5	2	+2
C	6	12	6	6	6	0
Mg	12	25	12	13	12	0
Tc	43	98	43	55	43	0
Ca	20	40	20	20	18	+2
Si	14	28	14	14	14	0
K	19	47	19	28	15	+4
Fe	26	56	26	30	23	+3
Br	35	79	35	44	36	-1
Ti	22	39	22	17	21	+1
P	15	30	15	15	15	0
Al	13	27	13	14	10	+3
S	16	32	16	16	16	0
Pd	46	106	46	60	45	+1
Cr	24	52	24	28	21	+3
Sn	50	118	50	68	50	0
Hg	80	200	80	120	79	+1
Au	79	197	79	118	78	+1

This page titled [1.2: Old Quantum Mechanics- Basic Developments](#) is shared under a [CC BY-NC 3.0](#) license and was authored, remixed, and/or curated by [Chris Schaller](#) via [source content](#) that was edited to the style and standards of the LibreTexts platform.