Chapter 29

Atomic Theory

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Atomic Theory

- Matter is composed of atoms
- Atoms are assembled from electrons, protons, and neutrons
- Atoms were discovered after Galileo, Newton, and Maxwell and most other physicists discussed so far had completed their work
- Quantum theory explains the way atoms are put together
- Key aspects of quantum theory explain the way atoms are put together
  - Can understand why different elements have different properties
  - Can explain the organization of the periodic table
Structure of the Atom

• By about 1890, most physicists and chemists believed matter was composed of atoms
• It was widely believed that atoms were indivisible
• Evidence for this picture of the atoms were the gas laws and the law of definite proportions
  • The law of definite proportions says that when a compound is completely broken down into its constituent elements, the masses of the constituent always have the same proportions
• It is now known that all the elements were composed of three different types of particles
  • Electrons, protons, and neutrons
Particle Review

- **Electrons**
  - Carry charge of \(-e\)
- **Protons**
  - Carry charge of \(+e\)
- **Neutrons**
  - Carry no net electrical charge
  - Mass approximately the same as the proton
- **Ideas of quantum theory need to be applied to understand the structure of the atom**
Plum Pudding Model

- Electrons were the first building-block particles to be discovered.
- The model suggested that the positive charge of the atom is distributed as a “pudding” with electrons suspended throughout the “pudding”.

Section 29.1
Plum Pudding Model, cont.

- A neutral atom has zero total electric charge
  - An atom must contain a precise amount of positive “pudding”
  - How was that accomplished?
- Physicists studied how atoms collide with other atomic-scale particles
  - Experiments carried out by Rutherford, Geiger and Marsden
Rutherford expected the positively charged pudding would have a low density
  • Since the charge was spread throughout the entire atom
• The alpha particle would pass freely through the “plum-pudding”
Planetary Model

- Most of the alpha particles did pass freely through the atom
- A small number of alpha particles were actually deflected through very large angles
  - Some bounced backward

Nucleus *not* drawn to scale. The nucleus is actually about 100,000 times smaller than the region occupied by electrons.

Section 29.1
• The reaction of the alpha particle could not be explained by the plum-pudding model
• Rutherford realized that all the positive charge in an atom must be concentrated in a very small volume
  • The mass and density of the positive charge was about the same as the alpha particle
• Most alpha particles missed this dense region and passed through the atom
• Occasionally an alpha particle collided with the dense region, giving it a large deflection
• He concluded that atoms contain a nucleus that is positively charged and has a mass much greater than that of the electron
• Rutherford suggested that the atom is a sort of miniature solar system

• The electrons orbit the nucleus just as the planets orbit the sun
  • The electrons must move in orbits to avoid falling into the nucleus as a result of the electric force

• The atomic nucleus contains protons
  • The charge on a proton is $+e$

• Since the total charge on an atom is zero, the number of protons must equal the number of electrons
Atomic Number and Neutrons

• The **atomic number** of the element is the number of protons it contains
  • Symbolized by Z
• Nuclei, except for hydrogen, also contain neutrons
• The neutron is a neutral particle
  • Zero net electric charge
• The neutron was discovered in the 1930s
• Protons are positively charged and repel each other
• The protons are attracted to the neutrons by an additional force that overcomes the Coulomb repulsion and holds the nucleus together

Section 29.1
Energy of Orbiting Electron

- The planetary model of the hydrogen atom is shown.
- Contains one proton and one electron.
- The electric force supplies a centripetal force.
- The speed of the electron is

\[ v = \sqrt{\frac{ke^2}{mr}} \]
Energy of Orbiting Electron, cont.

- This speed corresponds to a kinetic energy of the electron of $1.2 \times 10^{-18} \text{ J} = 7.5 \text{ eV}$
- This is the same order of magnitude as the measured ionization energy of the hydrogen atom of 13.6 eV
  - The ionization energy is the energy required to remove an electron from an atom in the gas phase
- The electron also has potential energy
  - The change in potential energy when the atom is ionized is $\sim 14 \text{ eV} $
Major Problem with the Planetary Model

- Stability of the electron orbit
  - Atoms are stable
  - Atoms should not be stable in the model
- There was no way to fix the planetary model to make the atom stable
Stability Details

- Since the electrons are undergoing accelerated motion, they should emit electromagnetic radiation.
- As the electron loses energy, it should spiral into inward to the nucleus.
- The atom would be inherently unstable.
- It should only last a fraction of a second.
Quantum Theory Solution

- Quantum theory avoids the problem of unstable electrons
  - It replaces orbits with discrete energy levels
- Quantum theory says the electrons are not simple particles that obey Newton’s laws and spiral into the nucleus
- The electron is a wave-particle described by a wave function with discrete energy levels
- Electrons gain or lose energy only when they undergo a transition between energy levels
Atomic Spectra

• The best evidence that an electron can exist only in discrete energy levels comes from the radiation an atom emits or absorbs when an electron undergoes a transition from one energy level to another.
• This was related to the question of what gives an object its color.
• Physicists of that time knew about the relationship between blackbody radiation and temperature.

Section 29.2
The sun’s spectrum shows sharp dips superimposed on the smooth blackbody curve.

The dips are called *lines* because of their appearance.

The dips show up as dark lines in the spectrum viewed by a prism or a diffraction grating.

The locations of the dips indicate the wavelengths at which the light intensity is lower than the expected blackbody value.

Section 29.2
Formation of Spectra

- When light from a pure blackbody source passes through a gas, atoms in the gas absorb light at certain wavelengths.
- The values of the wavelengths have been confirmed in the laboratory.
Absorption and Emission

• The dark spectral lines are called *absorption lines*
  • They result from the absorption of light
• The atoms can also produce an *emission spectrum*
  • The light is emitted by the atoms
• The absorption and emission lines occur at the same wavelengths
• The pattern of spectral lines is different for each element
Questions About Spectra

• Why do the lines occur at specific wavelengths?
• Why do absorption and emission lines occur at the same wavelength?
• What determines the pattern of wavelengths?
• Why are the wavelengths different for different elements?
Photon Energy

- The energy of a photon is $E_{\text{photon}} = h f$
- Since total energy is conserved, the energy of the photon is the difference in the energy of the atom before and after emission or absorption
- Since atomic emission occurs only at certain discrete wavelengths, the energy of the orbiting electron can only have certain discrete values
- According to Newton’s mechanics, the radius of the electron’s orbit can have a continuous range of values
Photon Energy and Newton

- Based on Newton’s mechanics, there is no way for the planetary orbit picture to give discrete electron energies
  - Energies of an orbiting electron could have a continuous range of values
  - So there is no way to explain the existence of discrete spectral lines
  - The problem is resolved in quantum mechanics’ explanation of the electron’s state in terms of a wave function instead of an orbit
Atomic Energy Levels

- The energy of an atom is quantized.
- The energy of an absorbed or emitted photon is equal to the difference in energy between two discrete atomic energy levels.
- The frequencies of the lines give the spacing between the atom’s energy levels.
- Explained the experimental evidence of discrete spectral lines.

Section 29.2
Bohr Model of the Atom

- Experiments showed that Rutherford’s planetary model of the atom did not work
  - Any model based on Newton’s mechanics would be a failure
- Bohr invented another model called the **Bohr model**
- Although not perfect, this model included ideas of quantum theory
  - Based on Rutherford’s planetary model
  - Included discrete energy levels
Ideas In Bohr’s Model

- Circular electron orbits
  - For simplification
- Used hydrogen
  - Simplest atom
- Postulated only certain electron orbits are allowed
  - To explain discrete spectral lines
  - Only specific values of r are allowed
  - Then only specific energies are allowed based on the values of r
- Energy level diagrams can be used to show absorption and emission of photons
- Explained the experimental evidence

Section 29.3
Energy Levels

- Each allowed orbit is a quantum state of the electron
- $E_1$ is the **ground state**
  - The state of lowest possible energy for the atom
- Other states are **excited states**
- Photons are emitted when electrons fall from higher to lower states
- When photons are absorbed, the electron undergoes a transition to a higher state

Section 29.3
Angular Momentum and \( r \)

- To determine the allowed values of the radius, \( r \), Bohr proposed that the orbital angular momentum of the electron could only have certain values
  
  \[ L = n \frac{h}{2\pi} \]

- \( n = 1, 2, 3, \ldots \) is an integer and \( h \) is Planck’s constant

- Combining this with the orbital motion of the electron, the radii of allowed orbits can be found
  
  \[ r = n^2 \left( \frac{h^2}{4\pi^2 mke^2} \right) \]
Values of $r$

- The only variable is $n$
  - The other terms in the equation for $r$ are constants
- The orbital radius of an electron in a hydrogen atom can have only these values
  - Shows the orbital radii are quantized
- The smallest value of $r$ corresponds to $n = 1$
  - This is called the *Bohr radius* of the hydrogen atom and is the smallest orbit allowed in the Bohr model
  - For $n = 1$, $r = 0.053$ nm
Energy Values

• The energies corresponding to the allowed values of \( r \) can also be calculated.

\[
E_{tot} = KE + PE_{elec} = -\left(\frac{2\pi^2 k^2 e^4 m}{h^2}\right) \frac{1}{n^2}
\]

• The only variable is \( n \), which is an integer and can have values \( n = 1, 2, 3, \ldots \).

• Therefore, the energy levels in the hydrogen atom are also quantized.

• For the hydrogen atom, this becomes

\[
E_{tot} = -\frac{13.6 \text{ eV}}{n^2}
\]
Energy Level Diagram for Hydrogen

- The negative energies come from the convention that $\text{PE}_{\text{elec}} = 0$ when the electron is infinitely far from the proton.
- The energy required to take the electron from the ground state and remove it from the atom is the ionization energy.
- The arrows show some possible transitions leading to emissions of photons.

\[ E_{\text{tot}} = -\frac{13.6}{n^2} \text{ eV} \]
Quantum Theory and the Kinetic Theory of Gases

- Quantum theory explains the claim that the collisions between atoms in a gas are elastic
- At room temperature, the kinetic energy of the colliding atoms is smaller than the spacing between the ground and the excited states
- A collision does not involve enough energy to cause a transition to a higher level
- The atoms stay in their ground state
  - None of their kinetic energy is converted into potential energy of the atomic electrons

Section 29.3
X-Rays from Atoms

- The highest photon energy available in a hydrogen atom is in the ultraviolet part of the electromagnetic spectrum.
- Other atoms can emit much more energetic photons.
- May applications use X-ray photons obtained from an electron transition from $E_2$ to $E_1$ in heavier atoms.
  - These are called $K_\alpha$ X-rays.
  - See table 29.1 for the energy of $K_\alpha$ X-rays produced by some elements.
Continuous Spectrum

- If an absorbed photon has more energy than is needed to ionize an atom, the extra energy goes into the kinetic energy of the ejected electron.
- This final energy can have a range of values and so the absorbed photon can have a range of values.
- This produces a continuous spectrum.

Section 29.3
• Bohr’s suggestion that the angular momentum of the electron is quantized was completely new.

• The assumption of quantized angular momentum can be understood in terms of de Broglie’s theory:
  • Which came about 10 years after Bohr made the assumption.

• De Broglie stated that electrons have a wave character, with a wavelength of \( \lambda = \frac{h}{p} \).
Bohr and de Broglie

- The allowed electron orbits in the Bohr model correspond to standing waves that fit into the orbital circumference.
- Since the circumference has to be an integer number of wavelengths, $2\pi r = n\lambda$.
- This leads to Bohr’s condition for angular momentum.

Section 29.3
Problems with Bohr’s Model

• The Bohr model was successful for atoms with one electron
  • H, He⁺, etc.
• The model does not correctly explain the properties of atoms or ions that contain two or more electrons
• Physicists concluded that the Bohr model is not the correct quantum theory
  • It was a “transition theory” that helped pave the way from Newton’s mechanics to modern quantum mechanics

Section 29.3
Modern Quantum Mechanics

- Modern quantum mechanics depends on the ideas of wave functions and probability densities instead of mechanical ideas of position and motion.
- To solve a problem in quantum mechanics, you use Schrödinger’s equations:
  - The solution gives the wave function, including its dependence on position and time.
- Four quantum numbers are required for a full description of the electron in an atom:
  - Bohr’s model used only one.
### TABLE 29.2 Quantum Numbers for Allowed Electron States in an Atom

<table>
<thead>
<tr>
<th>Quantum Number</th>
<th>Name</th>
<th>Possible Values</th>
</tr>
</thead>
<tbody>
<tr>
<td>( n )</td>
<td>Principal quantum number</td>
<td>( n = 1, 2, 3, \ldots )</td>
</tr>
<tr>
<td>( \ell )</td>
<td>Orbital quantum number</td>
<td>( \ell = 0, 1, 2, \ldots, n - 1 )</td>
</tr>
<tr>
<td>( m )</td>
<td>Orbital magnetic quantum number</td>
<td>( m = -\ell, -\ell + 1, \ldots, 0, \ldots, \ell - 1, \ell )</td>
</tr>
<tr>
<td>( s )</td>
<td>Spin quantum number</td>
<td>( s = -\frac{1}{2} ) or ( +\frac{1}{2} )</td>
</tr>
</tbody>
</table>
Principle Quantum Number

- $n$ is the principle quantum number
  - It can have values $n = 1, 2, 3, \ldots$
  - It is roughly similar to Bohr’s quantum number
  - As $n$ increases, the average distance from the electron to the nucleus increases
  - States with a particular value of $n$ are referred to as a “shell”
Orbital Quantum Number

• $\ell$ is the orbital quantum number
  • Allowed values are $\ell = 0, 1, 2, \ldots n - 1$
  • The angular momentum of the electron is proportional to $\ell$
    • States with $\ell = 0$ have no angular momentum
  • See the table for shorthand letters for various $\ell$ values

<table>
<thead>
<tr>
<th>Orbital Quantum Number</th>
<th>Configuration Letter</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\ell = 0$</td>
<td>$s$</td>
</tr>
<tr>
<td>$\ell = 1$</td>
<td>$p$</td>
</tr>
<tr>
<td>$\ell = 2$</td>
<td>$d$</td>
</tr>
<tr>
<td>$\ell = 3$</td>
<td>$f$</td>
</tr>
<tr>
<td>$\ell = 4$</td>
<td>$g$</td>
</tr>
<tr>
<td>$\ell = 5$</td>
<td>$h$</td>
</tr>
</tbody>
</table>

Section 29.4
Orbital Magnetic Quantum Number

- $m$ is the orbital magnetic quantum number
  - It has allowed values of $m = -\ell, -\ell + 1, \ldots, -1, 0, 1, \ldots, \ell$
  - You can think of $m$ as giving the direction of the angular momentum of the electron in a particular state
Spin Quantum Number

- \( s \) is the spin quantum number
  - \( s = + \frac{1}{2} \) or \( - \frac{1}{2} \)
    - These are often referred to as “spin up” and “spin down”
  - This gives the direction of the electron’s spin angular momentum
Electron Shells and Probabilities

- A particular quantized electron state is specified by all four of the quantum number $n, \ell, m$ and $s$
- The solution of Schrödinger’s equation also gives the wave function of each quantum state
  - From the wave function, you can calculate the probability for finding the electron at different locations around the nucleus
- Plots of probability distributions for an electron are often called “electron clouds”
Electron Clouds

The probability of finding an electron is largest where the “cloud” is darkest.

Ground state

1s
- \( n = 1 \)
- \( \ell = 0 \)
- \( m = 0 \)

2s
- \( n = 2 \)
- \( \ell = 0 \)
- \( m = 0 \)

2p
- \( n = 2 \)
- \( \ell = 1 \)
- \( m = 0 \)
- \( m = \pm 1 \)

A

B

\( p_x \) state

\( p_y \) state

\( p_z \) state
Electron Cloud Example

- Ground state of hydrogen
  - $n = 1$
  - The only allowed state for $\ell$ is $\ell = 0$
    - This is an “s state”
  - The only allowed state for $m$ is $m = 0$
  - The allowed states for $s$ are $s = \pm \frac{1}{2}$
    - The probability of finding an electron at a particular location does not depend on $s$, so both of these states have the same probability
  - The electron probability distribution forms a spherical “cloud” around the nucleus
    - See fig. 29.17 A
The electron probability distributions for all states are independent of the value of the spin quantum number.

For the hydrogen atom, the electron energy depends only on the value of $n$ and is independent of $\ell$, $m$ and $s$.

This is not true for atoms with more than one electron.
Multielectron Atoms

- The electron energy levels of multielectron atoms follow the same pattern as hydrogen
  - Use the same quantum numbers
- The electron distributions are also similar
- There are two main differences between hydrogen and multielectron atoms
  - The values of the electron energies are different for different atoms
  - The spatial extent of the electron probability clouds varies from element to element

Section 29.5
Pauli Exclusion Principle

- Each quantum state can be occupied by only one electron
  - Each electron must occupy its own quantum state, different from the states of all other electrons
- This is called the **Pauli exclusion principle**
- Each electron is described by a unique set of quantum numbers

Section 29.5
• The direction of the arrow represents the electron’s spin
• In C, the He electrons have different spins and obey the Pauli exclusion principle
Electron Configuration

- There is a useful shorthand notation for showing electron configurations

- Examples:
  - $1s^1$
    - $1 \rightarrow n = 1$
    - $s \rightarrow \ell = 0$
    - Superscript $1 \rightarrow 1$ electron
    - No information about electron spin
  - $1s^22s^22p^2$
    - 2 electrons in $n = 1$ with $\ell = 0$
    - 2 electrons in $n = 2$ with $\ell = 0$
    - 2 electrons in $n = 2$ with $\ell = 1$
The energy of each level depends mainly on the value of $n$.

In multielectron atoms, the order of energy levels is more complicated.

For shells higher than $n = 2$, the energies of subshells from different shells begin to overlap.

In general, the energy levels fill with electrons in the following order:

$1s\ 2s\ 2p\ 3s\ 3p\ 4s\ 3d\ 4p\ 5s\ 4d\ 5p\ 6s\ 4f$
Order of Energy Levels

Section 29.5
Chemical Properties of Elements

- Quantum theory explains the structure of the periodic table
- The periodic table was first assembled by Dmitry Mendelev in the late 1860’s
- Mendeleyev and other chemists had noticed that many elements could be grouped according to their chemical properties
- Mendeleyev organized his table by grouping related elements in the same column
- His table had a number of “holes” because many elements had not yet been discovered

Section 29.6
Mendeleyev could not explain why the regularities in the periodic table occurred.

The electron energy levels and the electron configuration of the atom are responsible for its chemical properties.

When an atom participates in a chemical reaction, some of its electrons combine with electrons from other atoms to form chemical bonds.

The bonding electrons are those occupying the highest energy levels.
# Periodic Table

<table>
<thead>
<tr>
<th>Symbol</th>
<th>Atomic Number</th>
<th>Atomic Mass</th>
<th>Configuration of outermost electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ca</td>
<td>20</td>
<td>40.078</td>
<td>4s(^2) 4p(^6) 5s(^2)</td>
</tr>
</tbody>
</table>

Note: Atomic mass values given are averaged over isotopes in the percentages in which they exist in nature.

For a description of the atomic data, visit physics.nist.gov/PhysRefData/Elements/per_text.html.

† For an unstable element, mass number of the most stable known isotope is given in parentheses.

†† Elements 114, 116, and 117 have not yet been named.

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**Section 29.6**
## Example Electron Configurations

### Table 29.4: Electron Configurations for Several Elements

<table>
<thead>
<tr>
<th>Element</th>
<th>Number of Electrons</th>
<th>Electron Configuration</th>
<th>Quantum Numbers of Occupied Electron Levels</th>
</tr>
</thead>
<tbody>
<tr>
<td>H (hydrogen)</td>
<td>1</td>
<td>1s^1</td>
<td>$n = 1, \ell = 0, m = 0, s = +\frac{1}{2}$</td>
</tr>
<tr>
<td>He (helium)</td>
<td>2</td>
<td>1s^2</td>
<td>$n = 1, \ell = 0, m = 0, s = +\frac{1}{2}$</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>$n = 1, \ell = 0, m = 0, s = -\frac{1}{2}$</td>
</tr>
<tr>
<td>Li (lithium)</td>
<td>3</td>
<td>1s^22s^1</td>
<td>Same as He plus</td>
</tr>
<tr>
<td>Be (beryllium)</td>
<td>4</td>
<td>1s^22s^2</td>
<td>$n = 2, \ell = 0, m = 0, s = +\frac{1}{2}$</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>$n = 2, \ell = 0, m = 0, s = -\frac{1}{2}$</td>
</tr>
<tr>
<td>B (boron)</td>
<td>5</td>
<td>1s^22s^22p^1</td>
<td>Same as Be plus</td>
</tr>
<tr>
<td>C (carbon)</td>
<td>6</td>
<td>1s^22s^22p^2</td>
<td>Same as Be plus</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>$n = 2, \ell = 1, m = 0, s = +\frac{1}{2}$</td>
</tr>
<tr>
<td>N (nitrogen)</td>
<td>7</td>
<td>1s^22s^22p^3</td>
<td>Same as C plus</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>$n = 2, \ell = 1, m = 1, s = +\frac{1}{2}$</td>
</tr>
<tr>
<td>O (oxygen)</td>
<td>8</td>
<td>1s^22s^22p^4</td>
<td>Same as C plus</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>$n = 2, \ell = 1, m = 1, s = -\frac{1}{2}$</td>
</tr>
<tr>
<td>F (fluorine)</td>
<td>9</td>
<td>1s^22s^22p^5</td>
<td>Same as O plus</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>$n = 2, \ell = 1, m = -1, s = +\frac{1}{2}$</td>
</tr>
<tr>
<td>Ne (neon)</td>
<td>10</td>
<td>1s^22s^22p^6</td>
<td>Same as O plus</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>$n = 2, \ell = 1, m = -1, s = +\frac{1}{2}$</td>
</tr>
<tr>
<td>Na (sodium)</td>
<td>11</td>
<td>1s^22s^22p^63s^1</td>
<td>Same as Ne plus</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>$n = 3, \ell = 0, m = 0, s = +\frac{1}{2}$</td>
</tr>
</tbody>
</table>
Electrons and Shells

• The electron that forms bonds with other atoms is a valence electron
• When a shell has all possible states filled it forms a closed shell
• Elements in the same column in the periodic table have the same number of valence electrons
• The last column in the periodic table contains elements with completely filled shells
  • These elements are largely inert
  • They almost never participate in chemical reactions

Section 29.6
Structure of the Periodic Table

- Mendeleyev grouped elements into *columns* according to their common bonding properties and chemical reactions
  - These properties rely on the valence electrons and can be traced to the electron configurations
- The *rows* correspond to different values of the principle quantum number, *n*
  - Since the *n* = 1 shell can hold only two electrons, the row contains only two elements
  - The number of elements in each row can be found by using the rules for allowed quantum numbers
Atomic Clocks

- Atomic clocks are used as global time standards
- The clocks are based on the accurate measurements of certain spectral line frequencies
- Cesium atoms are popular
- One second is now defined as the time it takes a cesium clock to complete 9,192,631,770 ticks

Section 29.7
Incandescent Light Bulbs

- The *incandescent bulb* contains a thin wire filament that carries a large electric current
  - Type developed by Edison
- The electrical energy dissipated in the filament heats it to a high temperature
- The filament then acts as a blackbody and emits radiation
Fluorescent Bulbs

- This type of bulb uses gas of atoms in a glass container
- An electric current is passed through the gas
- This produces ions and high-energy electrons
- The electrons, ions, and neutral atoms undergo many collisions, causing many of the atoms to be in an excited state
- These atoms decay back to their ground state and emit light
Neon and Fluorescent Bulbs

- A neon bulb contains a gas of Ne atoms
- Fluorescent bulbs often contain mercury atoms
  - Mercury emits strongly in the ultraviolet
  - The glass is coated with a fluorescent material
  - The photons emitted by the Hg atoms are absorbed by the fluorescent coating
  - The coating atoms are excited to higher energy levels
  - When the coating atoms undergo transitions to lower energy states, they emit new photons
  - The coating is designed to emit light throughout the visible spectrum, producing “white” light

Section 29.7
Lasers depend on the coherent emission of light by many atoms, all at the same frequency.

In *spontaneous emission*, each atom emits photons independently of the other atoms:
- It is impossible to predict when it will emit a photon.
- The photons are radiated randomly in all directions.

In a laser, an atom undergoes a transition and emits a photon in the presence of many other photons that have energies equal to the atom’s transition energy.

A process known as *stimulated emission* causes the light emitted by this atom to propagate in the same direction and with the same phase as surrounding light waves.
Laser is an acronym for light amplification by stimulated emission of radiation.

The light from a laser is thus a coherent source.

Mirrors are located at the ends of the bulb (laser tube).

One of the mirrors lets a small amount of the light pass through and leave the laser.
Lasers, final

- Laser can be made with a variety of different atoms
- One design uses a mixture of Ne and He gas and is called a helium-neon laser
  - The photons emitted by the He-Ne laser have a wavelength of about 633 nm
- Another common type of laser is based on light produced by light-emitting diodes (LEDs)
  - These photons have a wavelength around 650 nm
  - These are used in optical barcode scanners
Consider two hypothetical atoms and assume they are bound together to form a molecule. The binding energy of a molecule is the energy required to break the chemical bond between the two atoms. A typical bond energy is 10 eV.
• Assume the atom is pulled apart by separating the atoms a distance $\Delta x$

• The magnitude of the force between the atoms is

$$F = \left| \frac{\Delta PE}{\Delta x} \right|$$

• A $\Delta x$ of 1 nm should be enough to break the chemical bond

• This gives a force of $\sim 1.6 \times 10^{-19}$ N
Quantum Mechanics and Newtonian Mechanics

- Quantum mechanics is needed in the regime of electrons and atoms since Newton’s mechanics fails in that area.
- Newton’s laws work very well in the classical regime.
- Quantum theory can be applied to macroscopic objects, giving results that are virtually identical to Newton’s mechanics.
- Classical objects have extremely short wavelengths, making the quantum theory description in terms of particle-waves unnecessary.
Where the Regimes Meet

- Physicists are actively studying the area where quantum mechanics and Newtonian mechanics meet.
- Questions include:
  - The quantum behavior of living organisms such as viruses.
  - Wave function of the brain.