Electrons and Periodicity

History of Atomic Theory

Calculating Wavelength, Frequency, and Energy

History of the Periodic Table

Writing Electron Configurations

Valence Electrons and the Octet Rule

Periodic Trends
Democritus ≈ 400 BC

- Democritus was an ancient Greek philosopher.
- He was the first person to propose the idea that matter was composed of indivisible particles.
- He called these indivisible particles “atomos”.
John Dalton was an English school teacher.

He studied the work of scientists such as Lavoisier (Law of Conservation of Mass) and Proust (Law of Definite Proportions).
Dalton’s Atomic Theory

1. Each element is composed of extremely small particles called atoms.
2. All atoms of a given element are identical in size, mass, and other properties; atoms of different elements differ in size, mass, and other properties.
3. Atoms cannot be subdivided, created, or destroyed.
4. Atoms of different elements combine in simple whole-number ratios to form chemical compounds.
5. In chemical reactions, atoms are combined, separated or rearranged.
Dalton’s Atomic Theory

- At the time Dalton proposed his model of the atom, the subatomic particles had not been discovered.
- The discovery of electrons, protons and neutrons has led to modifications to Dalton’s atomic theory.

Which part(s) of Dalton’s atomic theory has(have) had to be modified and why?

Parts 2 and 3 have been modified. We now know that atoms are divisible into even smaller particles and that a given element can have atoms with different masses (isotopes).
J.J. Thomson

- Thomson experimented with cathode ray tubes and electricity.
- He concluded that electrons are a fundamental part of all atoms, regardless of the element.
- Thomson measured the ratio of the charge to mass of an electron.
J.J Thomson’s Plum Pudding Model of the Atom

Thomson’s model consisted of a spherically shaped atom composed of a uniformly distributed positive charge within which the individual negatively charged electrons resided.
Robert Millikan succeeded in measuring the **charge** and **mass** of an electron by performing what is known as the “Millikan oil-drop experiment”.

In his experiment small drops of oil, which had picked up extra electrons, were allowed to fall between two electrically charged plates.

Millikan monitored the drops, measuring how the voltage on the plates affected their rate of fall. From these data he calculated the charge and mass of the electron.
Ernest Rutherford was interested in studying how positively charged alpha particles interacted with solid matter.

He and his associates, Hans Geiger and Ernest Marsden, conducted an experiment where they bombarded a thin sheet of gold foil with alpha particles.
Experimental Set-up for Gold Foil Experiment
What did Rutherford and his associates hypothesize would happen?

- Based upon the Plum Pudding Model of the atom it was assumed that most of the mass and charge of the atom were uniformly distributed throughout the gold foil.

- Rutherford and his associates hypothesized that the alpha particles would pass through the gold foil with only a slight deflection.
Most of the alpha particles did pass through with only a slight deflection.

A few of the particles bounced back towards the source.
Rutherford concluded that most of the volume of the atom was empty space. Why did he make this conclusion?

Most of the alpha particles were able to go straight through with little or no deflection implying that they had not hit anything.
Rutherford also concluded that the atom contained a small, dense, positively charged nucleus. Why did he make this conclusion?

Only a few of the positively charged alpha particles were widely deflected or bounced back. These particles must have encountered something small and positively charged. (like charges repel)
Rutherford’s Nuclear Model of the Atom

Rutherford suggested that the electrons surrounded the positively charged nucleus.
Rutherford’s Prediction

Since the atom was more massive than the mass of just the protons and electrons, Rutherford predicted the existence of neutral particles which we now call neutrons.
Doubts about Rutherford’s Model of the Atom

Rutherford’s model of the atom seemed to defy the laws of physics.

There was no explanations as to why the electrons did not continuously give off energy as they accelerated around the nucleus and eventually collapse into the nucleus.
In 1932 Chadwick confirmed the existence of neutrons through his experimentation with isotopes and radiation.
Niels Bohr attempted to solve the mystery as to what kept the electron in orbit.

He based his model of the atom on the atomic emission spectrum of hydrogen.

The Bohr Model of the atoms is often called the “planetary model”.
Atomic Emission Spectrum for Hydrogen

Continuous Spectrum

Atomic Emission Spectrum for Hydrogen
1. Bohr explained that the electrons in an atom exist in **specific energy levels** and they cannot be found anywhere between energy levels.

2. Unlike the current beliefs in classical physics, Bohr believed that the energy changes were not **continuous** and that electrons could not move from one energy level to another unless they absorbed or released the specific amount of energy associated with that energy change.

3. According to the Bohr model of the atom, hydrogen’s atomic emission spectrum results from electrons falling from **higher-energy** atomic orbits to **lower-energy** atomic orbits.

4. Unfortunately, the Bohr model of the atom only worked for the **hydrogen** atom.
Quantum Mechanical Model of the Atom

The current model of the atom is known as the **Quantum Mechanical Model**.

The work of many different scientists led to the quantum mechanical model.
Max Planck

Planck determined that electromagnetic energy is **quantized**.

That is, for a given frequency of radiation (or light), all possible energies are **multiples** of a certain unit of energy called a quantum.

\[ E = h\nu \]
In 1924, de Broglie suggested that, like light, electrons could as both particles and waves.

This is known as wave-particle duality.
Werner Heisenberg

Heisenberg is best known for discovering one of the central principles of modern physics, the **Heisenberg uncertainty principle**, and for the development of quantum mechanics.

The Heisenberg uncertainty principle states that it is impossible to determine simultaneously both the **position** and **momentum** of an electron.
Schrödinger is famous for his wave equation.

His wave equation treated the hydrogen’s atom as a wave. Unlike the Bohr model, Schrödinger’s new model for the hydrogen atom seemed to apply equally well to atoms of other elements.
Quantum Mechanical Model

Unlike the Bohr model, the quantum mechanical model does not define the exact path an electron takes around the nucleus.

It is concerned with the probability of finding an electron in a certain position.

This probability can be portrayed as a blurry cloud of negative charge.

The cloud is the most dense where the probability of finding an electron is large.
What is Light?

Light is *electromagnetic radiation* that is visible to our eyes.
Electromagnetic Spectrum

- **Increasing Energy**
- **Increasing Wavelength**

<table>
<thead>
<tr>
<th></th>
<th>0.0001 nm</th>
<th>0.01 nm</th>
<th>10 nm</th>
<th>1000 nm</th>
<th>0.01 cm</th>
<th>1 cm</th>
<th>1 m</th>
<th>100 m</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Gamma rays</strong></td>
<td>X-rays</td>
<td><strong>Ultra-violet</strong></td>
<td><strong>Infrared</strong></td>
<td><strong>Radio waves</strong></td>
<td>Radar TV FM AM</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

- **Visible Light**
  - 400 nm to 700 nm
Electromagnetic Spectrum
Wave-Particle Duality

During the early 1900’s, scientist came to the realization that light cannot be defined exclusively as a wave or as a particle, because it displays characteristics of both.
Light as a Wave

Light exhibits wave properties when it is reflected off a surface.
Light as a Wave

Light exhibits wave properties when it is bent (refracted) as it passes between material of different optical density.
Light as a Wave

Light exhibits wave properties when it is **diffracted** as it bends around obstacles in its path.
Light as a Wave

Light exhibits wave properties when it exhibits interference.
Properties of Waves:

- \( \lambda \) Wave length – distance from crest to crest.
- \( c \) Speed of light in a vacuum = \( 3.00 \times 10^8 \) m/s.
- \( T \) Period – Time between passage of successive crests.
- \( \nu \) Frequency – Number of crest passages per unit time.
- \( A \) Amplitude – Distance from level of crest to level of trough.
Calculating Frequency and Wavelength

The frequency and wavelength of a wave are inversely proportional.

\[ \lambda = \frac{c}{\nu} \]

- **c**: speed of light (3.00 \( \times \) 10\(^8\) m/s)
- **\( \lambda \)**: wavelength (m, nm, etc.)
- **\( \nu \)**: frequency (Hz)
Find the frequency of a photon with a wavelength of 434 nm. 
(1 nm = 1\times10^{-9} \text{ m})

\[ \lambda=434 \text{ nm} = 4.34\times10^{-7} \text{ m} \]

\[ \nu = \frac{c}{\lambda} \quad \nu = \frac{3.00 \times 10^8 \text{ m/s}}{4.34 \times 10^{-7} \text{ m}} \]

\[ \nu = 6.91\times10^{14} \text{ Hz} \]
Light as a Particle

Light exhibits particle properties when it demonstrates the **photoelectric effect**.

The photoelectric effect is the ejection of electrons from a metal surface as a result of the exposure to high energy light.
The specific particle of light energy that can be emitted or absorbed as electromagnetic radiation is called a photon.

The packet of energy carried by the photon is called a quantum of energy.

Since the light energy is available only in discrete amounts, it is said to be quantized.
Calculating Energy Changes

The amount of energy released or absorbed by a photon can be calculated using the following equation.

\[ E = h \nu \]

- **E** = energy (J)
- **h** = Planck’s constant = \(6.626 \times 10^{-34}\) J \(\cdot\) s
- **\nu** = frequency (Hz or s\(^{-1}\))
What is the energy of a photon that has a frequency of $6.32 \times 10^{20}$ Hz?

$$E = h \nu$$

$$E = (6.626 \times 10^{-34} \text{ J} \cdot \text{s}) (6.32 \times 10^{20} \text{ Hz})$$

$$E = 4.19 \times 10^{-13} \text{ J}$$
Mathematical Relationships Regarding Electromagnetic Radiation

The equations

\[ \lambda = \frac{c}{\nu} \quad \text{and} \quad E = h\nu \]

can be combined to form the following equation:

\[ E = \frac{hc}{\lambda} \]
The blue color in fireworks is often achieved by heating copper(I) chloride (CuCl) to about 1200°C. Then the compound emits blue light having a wavelength of 450. nm. What is the increment of energy (the quantum) that is emitted at 450.nm by CuCl?

450. nm = $4.50 \times 10^{-7}$ m

$$E = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34} \text{J} \cdot \text{s})(3.00 \times 10^8 \text{m})}{(4.50 \times 10^{-7} \text{m})} = 4.42 \times 10^{-19} \text{J}$$
Absorption and Emission of Energy

When electrons absorb energy, the electrons move from lower energy levels to higher levels.

When an electron falls from a higher energy level to a lower energy level, energy is emitted as electromagnetic radiation.

The frequency of the radiation emitted or absorbed by the electron depends upon the energy difference between the levels.
Spectroscopy

Spectroscopy is the study of the colors of light emitted or absorbed by atoms and molecules.

Spectroscopy can be used to identify the composition of stars and other stellar material.
The **frequency** of light, or the number of vibrations per second, determines the **color** of the light.

If white light is passed through a prism or diffraction grating, the colors of light are spread out into a **continuous spectrum**.
If the light produced by an element is passed through a prism or diffraction grating, a bright line (emission) spectrum will be produced.

Atomic emission spectra are often thought of as fingerprints of the elements.
If white light is passed through a chemical substance and then passed through a prism or diffraction grating, an absorption spectrum will be produced.

Absorption spectrum for hydrogen.
Comparison of Spectrum

Continuous Spectrum

Bright Line (Emission) Spectrum

Absorption Spectrum
Lab: Atomic Spectra
History of the Periodic Table

Scientists throughout the years have attempted to identify ways in which the elements could be organized.
In 1817, Johann Döbereiner grouped elements into sets of 3 based on their properties and called them triads.
Döbereiner’s Law of Triads

The elements in a triad had similar chemical properties. The physical properties varied by atomic mass.

<table>
<thead>
<tr>
<th>Element</th>
<th>Atomic Mass (amu)</th>
<th>Density (g/mL)</th>
<th>Melting Point (°C)</th>
<th>Boiling Point (°C)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Chlorine</td>
<td>35.5</td>
<td>0.00321</td>
<td>-101</td>
<td>-34</td>
</tr>
<tr>
<td>Bromine</td>
<td>79.9</td>
<td>3.12</td>
<td>-7</td>
<td>59</td>
</tr>
<tr>
<td>Iodine</td>
<td>127</td>
<td>4.93</td>
<td>114</td>
<td>185</td>
</tr>
</tbody>
</table>

The atomic mass of the middle element of the triad was approximately equal to the average of the other two elements. Unfortunately not all of the known elements fit into triads.
Newlands’s Law of Octaves

John Newlands classified the 49 known elements in order of increasing atomic mass.
Newlands’s Octaves

Newlands found that the elements seemed to have recurring similarities of properties every 8th element.

Unfortunately, elements that were later discovered (i.e. the noble gases) did not fit into his pattern.
Mendelev’s Periodic Table

In 1869, Mendeleev created a periodic table in which he listed the 63 known elements in order of increasing atomic mass.
Mendeleev’s Periodic Table

Mendeleev grouped elements with similar chemical properties together and left blank spaces (gaps) where there were no known elements with the appropriate properties and masses.
Mendeleev’s Periodic Table

Mendeleev grouped elements with similar chemical properties together and left blank spaces (gaps) where there were no known elements with the appropriate properties and masses.

<table>
<thead>
<tr>
<th>Element</th>
<th>Atomic Number</th>
<th>Mass Number</th>
<th>Symbol</th>
<th>Properties</th>
</tr>
</thead>
<tbody>
<tr>
<td>Be</td>
<td>4</td>
<td>9.4</td>
<td>Be</td>
<td></td>
</tr>
<tr>
<td>Mg</td>
<td>12</td>
<td>24</td>
<td>Mg</td>
<td></td>
</tr>
<tr>
<td>Zn</td>
<td>30</td>
<td>65.39</td>
<td>Zn</td>
<td></td>
</tr>
<tr>
<td>Cd</td>
<td>48</td>
<td>112.41</td>
<td>Cd</td>
<td></td>
</tr>
<tr>
<td>B</td>
<td>5</td>
<td>11</td>
<td>B</td>
<td></td>
</tr>
<tr>
<td>Al</td>
<td>13</td>
<td>27.0</td>
<td>Al</td>
<td></td>
</tr>
<tr>
<td>Si</td>
<td>14</td>
<td>28.09</td>
<td>Si</td>
<td></td>
</tr>
<tr>
<td>C</td>
<td>6</td>
<td>12.01</td>
<td>C</td>
<td></td>
</tr>
<tr>
<td>N</td>
<td>7</td>
<td>14.002</td>
<td>N</td>
<td></td>
</tr>
<tr>
<td>O</td>
<td>8</td>
<td>16.0000</td>
<td>O</td>
<td></td>
</tr>
<tr>
<td>F</td>
<td>9</td>
<td>19.00</td>
<td>F</td>
<td></td>
</tr>
<tr>
<td>Na</td>
<td>11</td>
<td>22.99</td>
<td>Na</td>
<td></td>
</tr>
<tr>
<td>K</td>
<td>19</td>
<td>39.098</td>
<td>K</td>
<td></td>
</tr>
<tr>
<td>Rb</td>
<td>37</td>
<td>85.47</td>
<td>Rb</td>
<td></td>
</tr>
<tr>
<td>Cs</td>
<td>55</td>
<td>133.34</td>
<td>Cs</td>
<td></td>
</tr>
<tr>
<td>Ba</td>
<td>56</td>
<td>137.336</td>
<td>Ba</td>
<td></td>
</tr>
<tr>
<td>La</td>
<td>57</td>
<td>138.907</td>
<td>La</td>
<td></td>
</tr>
<tr>
<td>Ce</td>
<td>58</td>
<td>140.12</td>
<td>Ce</td>
<td></td>
</tr>
<tr>
<td>Pr</td>
<td>59</td>
<td>140.907</td>
<td>Pr</td>
<td></td>
</tr>
<tr>
<td>Nd</td>
<td>60</td>
<td>144.24</td>
<td>Nd</td>
<td></td>
</tr>
<tr>
<td>Pm</td>
<td>61</td>
<td>145.00</td>
<td>Pm</td>
<td></td>
</tr>
<tr>
<td>Sm</td>
<td>62</td>
<td>150.36</td>
<td>Sm</td>
<td></td>
</tr>
<tr>
<td>Eu</td>
<td>63</td>
<td>151.96</td>
<td>Eu</td>
<td></td>
</tr>
<tr>
<td>Gd</td>
<td>64</td>
<td>157.25</td>
<td>Gd</td>
<td></td>
</tr>
<tr>
<td>Tb</td>
<td>65</td>
<td>158.93</td>
<td>Tb</td>
<td></td>
</tr>
<tr>
<td>Dy</td>
<td>66</td>
<td>162.50</td>
<td>Dy</td>
<td></td>
</tr>
<tr>
<td>Ho</td>
<td>67</td>
<td>164.93</td>
<td>Ho</td>
<td></td>
</tr>
<tr>
<td>Er</td>
<td>68</td>
<td>167.26</td>
<td>Er</td>
<td></td>
</tr>
<tr>
<td>Tm</td>
<td>69</td>
<td>168.93</td>
<td>Tm</td>
<td></td>
</tr>
<tr>
<td>Yb</td>
<td>70</td>
<td>173.04</td>
<td>Yb</td>
<td></td>
</tr>
<tr>
<td>Lu</td>
<td>71</td>
<td>174.97</td>
<td>Lu</td>
<td></td>
</tr>
<tr>
<td>Hf</td>
<td>72</td>
<td>178.49</td>
<td>Hf</td>
<td></td>
</tr>
<tr>
<td>Ta</td>
<td>73</td>
<td>180.95</td>
<td>Ta</td>
<td></td>
</tr>
<tr>
<td>W</td>
<td>74</td>
<td>183.84</td>
<td>W</td>
<td></td>
</tr>
<tr>
<td>Re</td>
<td>75</td>
<td>186.21</td>
<td>Re</td>
<td></td>
</tr>
<tr>
<td>Os</td>
<td>76</td>
<td>190.23</td>
<td>Os</td>
<td></td>
</tr>
<tr>
<td>Ir</td>
<td>77</td>
<td>192.22</td>
<td>Ir</td>
<td></td>
</tr>
<tr>
<td>Pt</td>
<td>78</td>
<td>195.08</td>
<td>Pt</td>
<td></td>
</tr>
<tr>
<td>Au</td>
<td>79</td>
<td>196.97</td>
<td>Au</td>
<td></td>
</tr>
<tr>
<td>Hg</td>
<td>80</td>
<td>200.59</td>
<td>Hg</td>
<td></td>
</tr>
<tr>
<td>Tl</td>
<td>81</td>
<td>204.38</td>
<td>Tl</td>
<td></td>
</tr>
<tr>
<td>Pb</td>
<td>82</td>
<td>207.20</td>
<td>Pb</td>
<td></td>
</tr>
<tr>
<td>Bi</td>
<td>83</td>
<td>208.98</td>
<td>Bi</td>
<td></td>
</tr>
<tr>
<td>Po</td>
<td>84</td>
<td>209.00</td>
<td>Po</td>
<td></td>
</tr>
<tr>
<td>At</td>
<td>85</td>
<td>210.00</td>
<td>At</td>
<td></td>
</tr>
<tr>
<td>Rn</td>
<td>86</td>
<td>222.00</td>
<td>Rn</td>
<td></td>
</tr>
<tr>
<td>Fr</td>
<td>87</td>
<td>223.00</td>
<td>Fr</td>
<td></td>
</tr>
<tr>
<td>Ra</td>
<td>88</td>
<td>226.02</td>
<td>Ra</td>
<td></td>
</tr>
<tr>
<td>Ac</td>
<td>89</td>
<td>227.03</td>
<td>Ac</td>
<td></td>
</tr>
<tr>
<td>Th</td>
<td>90</td>
<td>232.03</td>
<td>Th</td>
<td></td>
</tr>
<tr>
<td>Pa</td>
<td>91</td>
<td>231.03</td>
<td>Pa</td>
<td></td>
</tr>
<tr>
<td>U</td>
<td>92</td>
<td>238.03</td>
<td>U</td>
<td></td>
</tr>
<tr>
<td>Np</td>
<td>93</td>
<td>237.05</td>
<td>Np</td>
<td></td>
</tr>
<tr>
<td>Pu</td>
<td>94</td>
<td>244.09</td>
<td>Pu</td>
<td></td>
</tr>
<tr>
<td>Am</td>
<td>95</td>
<td>243.06</td>
<td>Am</td>
<td></td>
</tr>
<tr>
<td>Cm</td>
<td>96</td>
<td>247.07</td>
<td>Cm</td>
<td></td>
</tr>
<tr>
<td>Bk</td>
<td>97</td>
<td>247.07</td>
<td>Bk</td>
<td></td>
</tr>
<tr>
<td>Cf</td>
<td>98</td>
<td>251.04</td>
<td>Cf</td>
<td></td>
</tr>
<tr>
<td>Es</td>
<td>99</td>
<td>252.04</td>
<td>Es</td>
<td></td>
</tr>
<tr>
<td>Fm</td>
<td>100</td>
<td>257.04</td>
<td>Fm</td>
<td></td>
</tr>
<tr>
<td>Md</td>
<td>101</td>
<td>258.04</td>
<td>Md</td>
<td></td>
</tr>
<tr>
<td>Corp</td>
<td>102</td>
<td>261.04</td>
<td>Corp</td>
<td></td>
</tr>
<tr>
<td>No</td>
<td>103</td>
<td>262.04</td>
<td>No</td>
<td></td>
</tr>
<tr>
<td>Lr</td>
<td>104</td>
<td>266.04</td>
<td>Lr</td>
<td></td>
</tr>
<tr>
<td>Rf</td>
<td>105</td>
<td>267.04</td>
<td>Rf</td>
<td></td>
</tr>
<tr>
<td>Db</td>
<td>106</td>
<td>276.04</td>
<td>Db</td>
<td></td>
</tr>
<tr>
<td>Sg</td>
<td>107</td>
<td>278.04</td>
<td>Sg</td>
<td></td>
</tr>
<tr>
<td>Bh</td>
<td>108</td>
<td>272.04</td>
<td>Bh</td>
<td></td>
</tr>
<tr>
<td>Hs</td>
<td>109</td>
<td>277.04</td>
<td>Hs</td>
<td></td>
</tr>
<tr>
<td>Mt</td>
<td>110</td>
<td>278.04</td>
<td>Mt</td>
<td></td>
</tr>
<tr>
<td>Ds</td>
<td>111</td>
<td>281.04</td>
<td>Ds</td>
<td></td>
</tr>
<tr>
<td>Rg</td>
<td>112</td>
<td>286.04</td>
<td>Rg</td>
<td></td>
</tr>
</tbody>
</table>

Д. Менделеев
Mendeleev’s Periodic Table

Mendeleev was able to predict the physical and chemical properties of many of the missing elements.
One of the elements Mendeleev predicted the properties of was **Eka-boron**.

Eka-boron was located between calcium and titanium on the periodic table.

What is the name of eka-boron now? **Scandium**
Mendeleev’s Periodic Table

Mendeleev stated the following periodic law: The properties of the elements are a periodic function of their atomic masses.

However, not all of the elements fit into the periodic table in order of increasing atomic mass.

Mendeleev arranged tellurium and iodine and cobalt and nickel out of order by atomic mass so that they could be placed in the groups with which they shared similar chemical properties.

(Mendeleev believed that the atomic masses of tellurium and iodine and cobalt and nickel had been incorrectly determined.)
In 1913, Henry Moseley used x-ray diffraction to determine the number of protons in the nucleus of an atom.

The periodic table was then arranged in order of increasing atomic number.
Interesting Fact about Henry Moseley

In 1914 Henry Moseley enlisted in the British armed forces.

He was fatally shot during the Battle of Gallipoli at the age of 27.

Because of Moseley's death in World War I, the British government instituted a policy of not allowing its prominent and promising scientists to enlist for combat duty in the armed forces of the Crown.
The Modern Periodic Law

The Modern Periodic Law states: The physical and chemical properties of the elements are periodic functions of their atomic numbers.
In 1944, Seaborg rearranged the structure of the periodic table based on the properties of elements he discovered. He removed Th, Pa, and U from the body of the table. The elements became the beginning of the actinide series.

Seaborg’s new arrangement allowed him to predict the properties of even more elements. He was the principal or co-discover of plutonium, americium, curium, berkelium, californium, einsteinium, fermium, mendelevium, nobelium and seaborgium.
Seaborg was part of the Manhattan Project. He and a group of scientists were instrumental in discovering (creating) plutonium from uranium. Plutonium was used in the second atomic bomb.

He made many contributions to nuclear medicine. He developed numerous isotopes of elements with important applications in the diagnosis and treatment of diseases, most notably iodine-131, which is used in the treatment of thyroid disease.

Element 106 was named Seaborgium (Sg) after him. He was still alive at the time it was named.
Arrangement of the Periodic Table

What do you already know about the periodic table?
Arrangement of the Periodic Table

1. Columns of elements in the periodic table are called **groups** or **families**.

2. There are **18** groups of elements.

3. The elements within the same family have **similar**, but not identical chemical properties.

4. The representative groups are groups **1,2** and **13-18**. These were previously called the “**A”** groups on the periodic table.

5. Elements found within those groups are called **representative** elements or **main group** elements.
Arrangement of the Periodic Table

6. Each row of elements in the periodic table is called a period.

7. There are 7 periods of elements.

8. The elements in a period are not alike in properties.

9. The first element in a period is generally a very active metal. The last element in a period is always a very inactive gas.
Properties of Metals, Nonmetals and Metalloids
Properties of Metals
(located on the left side of the periodic table)

1. They are malleable, ductile and have luster.
2. They reflect heat, light, and other forms of electromagnetic radiation.
3. They are good conductors of heat and electricity.
4. They have relatively high density.
5. They are solids at room temperature, except for mercury, which is a liquid.
6. They have relatively high melting points, except for mercury and gallium.
7. They don’t combine chemically with other metals. Metals combine physically to create alloys. Examples of alloys include: brass, stainless steel and solder.
8. They lose electrons in chemical reactions to become cations (positive ions).
9. They react with acids to produce hydrogen gas.
Properties of Nonmetals
(located on the right side of the periodic table)

1. They are dull and brittle.
2. They don’t conduct heat and electricity well.
3. They have relatively low boiling and freezing points.
4. They exist in all three phases at room temperature, but most are gases.
5. They gain electrons in chemical reactions to become anions (negative ions), except for the noble gases.
6. Many nonmetals are diatomic. This means they exist as molecules of two atoms. The seven common diatomic elements are H₂, Br₂, O₂, N₂, Cl₂, I₂, and F₂. (You need to memorize the list of diatomic elements.)
Properties of Metalloids
(located along the stair step on the periodic table)

1. They possess intermediate properties between metals and nonmetals.
2. They tend to be brittle.
3. They are semiconductors at higher than room temperatures.
4. They are all solids at STP.

antimony  Germanium  Silicon
Special Groups in the Periodic Table

Periodic Table of Elements

- Noble Gases
- Inner Transition Metals
- Metals

Legend - click to find out more...

- H - gas
- Li - solid
- Br - liquid
- Tc - synthetic

- Non-Metals
- Transition Metals
- Inner Transition Metals
- Halogens
- Alkali Metals
- Alkali Earth Metals
- Other Metals
- Noble Gases
Alkali Metals

1. They are soft and easily cut.
2. They have relatively low densities which increase going down the group. They have relatively low boiling and freezing points which decrease going down the group.
3. They are highly reactive. Reactivity increases going down the group.
4. They are usually found combined with elements because they are highly reactive.
5. They react chemically with H₂O, Cl₂, F₂, and O₂.
6. They oxidize rapidly in air and lose luster quickly.
7. They form basic hydroxides in chemical reactions with water.
8. They react with acids to produce hydrogen gas.
9. They give off distinct colors when exposed to flame: Na - yellow, Li - red, K - violet, Rb - maroon, and Cs - magenta.
10. Hydrogen is a nonmetal, but it is placed in this group because it combines with other elements similarly to the elements in this group.
### Alkaline Earth Metals

1. They are not as soft as alkali metals.
2. They are denser than the alkali metals.
3. They have higher boiling and freezing points than the alkali metals.
4. They are less reactive than the alkali metals.
5. They react chemically with oxygen to form basic oxides.
6. They react chemically with hot or boiling water and steam to form basic oxides.
7. They react chemically with acids to produce hydrogen gas.
Halogens

1. This group includes solids, liquids, and gases at room temperature.
2. They are poor conductors of heat and electricity.
3. They have boiling and freezing points that increase going down the group.
4. They are highly reactive and react violently with hot metals.
5. The word halogen means “salt former”. They form white crystalline salts with alkali metals.
6. They form diatomic molecules that require high temperatures to break bonds.
7. They are good oxidizers and react vigorously with organic compounds.
Noble Gases

1. They are colorless and odorless.
2. They have very low boiling and freezing points.
3. They exist as single atoms and rarely combine with other elements (nonreactive chemically).
4. Compounds containing noble gases are generally unstable.
Transition Metals

<table>
<thead>
<tr>
<th></th>
<th>3</th>
<th>4</th>
<th>5</th>
<th>6</th>
<th>7</th>
<th>8</th>
<th>9</th>
<th>10</th>
<th>11</th>
<th>12</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sc</td>
<td>Ti</td>
<td>V</td>
<td>Cr</td>
<td>Mn</td>
<td>Fe</td>
<td>Co</td>
<td>Ni</td>
<td>Cu</td>
<td>Zn</td>
<td></td>
</tr>
<tr>
<td>Y</td>
<td>Zr</td>
<td>Nb</td>
<td>Mo</td>
<td>Tc</td>
<td>Ru</td>
<td>Rh</td>
<td>Pd</td>
<td>Ag</td>
<td>Cd</td>
<td></td>
</tr>
<tr>
<td>La</td>
<td>Hf</td>
<td>Ta</td>
<td>W</td>
<td>Re</td>
<td>Os</td>
<td>Ir</td>
<td>Pt</td>
<td>Au</td>
<td>Hg</td>
<td></td>
</tr>
<tr>
<td>Ac</td>
<td>Rb</td>
<td>Db</td>
<td>Sg</td>
<td>Bh</td>
<td>Hs</td>
<td>Mt</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

1. They possess characteristics of active metals to varying degrees.
2. Most transition elements have high freezing and boiling points.
3. They are generally denser than alkali metals and alkaline earth metals.
4. They form several different compounds with a given nonmetal.
5. They form compounds that are usually brightly colored.
1. They are sometimes called the “Rare Earth Elements”
2. They include the lanthanide and actinide series.
3. They are very dense compared to most other metals.
4. Many of the inner transition metals are radioactive.
5. Elements past uranium are called “transuranium” elements and they are synthetic.
Quantum Mechanical Model vs. the Bohr Model

The Bohr model of the atom was an early quantum mechanical model.

The quantum mechanical model, like the Bohr model, restricts the energy of electrons to certain values.

Unlike the Bohr model, the quantum mechanical model does not define an exact path an electron takes around the nucleus.

The Heisenberg uncertainty principal states that it is impossible to know simultaneously both the velocity and the position of an electron. (This is due to the fact that electrons have both wave and particle like behavior.)
Where are the electrons located in the quantum mechanical model?

In the quantum mechanical model of the atom, the probability of finding an electron within a certain volume of space surrounding the nucleus can be represented as a fuzzy cloud.

The cloud is more dense where the probability of finding the electron is high.

An atomic orbital is a region in space where there is a high probability of finding an electron.
Electron Configurations and Orbital Diagrams

Electron configurations and orbital diagrams can be used to describe the way in which electrons are arranged around the nuclei of atoms.

Three rules – the aufbau principle, the Pauli Exclusion principle, and Hund’s rule – must be followed when writing electron configurations and orbital diagrams.
Aufbau Principle

The aufbau principle states that electrons enter orbitals of lowest energy first.

Energy levels of electrons are represented by the principal quantum number (n).

The principal energy levels are assigned values in order of increasing energy: n=1, 2, 3, 4, and so fourth.

The average distance of the electron from the nucleus increases with increasing values of n.
Sublevels

Within each principal energy level, the electrons occupy energy sublevels.

There are four types of sublevels: s, p, d, and f.

Each sublevel is made of atomic orbitals.

The s sublevel contains 1 orbital.
The p sublevel contains 3 orbitals.
The d sublevel contains 5 orbitals.
The f sublevel contains 7 orbitals.
Illustration of the s and p orbitals

- **s Orbital**
- **$p_x$ Orbital**
- **$p_y$ Orbital**
- **$p_z$ Orbital**
- **All $p$ orbitals full**
Aufbau Principle

Here is a diagram illustrating the order in which the electrons enter the orbitals.
Pauli Exclusion Principle

The Pauli exclusion principal states that only two electrons can occupy an atomic orbital and that they must have opposite spin.

Based on the Pauli exclusion principle, how many electrons can occupy

- the s sublevel? 2
- the p sublevel? 6
- the d sublevel? 10
- the f sublevel? 14
Hund’s Rule

When electrons occupy orbitals of equal energy, one electron occupies each orbital until all orbitals contain one electron with parallel spins.
Example 1 - Carbon

First determine the number of electrons.

Carbon has 6 electrons.

We will use boxes to represent the atomic orbitals and arrows to represent the electrons.

The electron configuration would be:

\[ 1s^2 2s^2 2p^2 \]
Example 2 - Selenium

First determine the number of electrons.

Selenium has 34 electrons.

The electron configuration would be:

$$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^4$$
You Try It

1. Write orbital diagrams for the following elements.
   a. Fluorine
   b. Calcium
   c. Arsenic
You Try It

2. Write electron configurations for the following elements.
   a. Magnesium
   b. Sulfur
   c. Iodine
You Try It

3. Which elements are represented by each of the following electron configurations?

   a. \(1s^22s^22p^4\)

   b. \(1s^22s^22p^63s^23p^64s^23d^{10}4p^4\)

   c. \(1s^22s^22p^63s^23p^2\)
The electron configurations and orbital diagrams we have been drawing represent atoms in the ground state. The ground state is the lowest energy state of an atom.

When energy is absorbed by an atom, the electrons move to an excited state.

Here is an example of an electron configuration for an atom in an excited state: $1s^22s^22p^53s^1$
Electron Configurations and the Periodic Table

With a few exceptions, correct electron structures for atoms can be derived from examining the element’s position on the periodic table.
How are electron configurations related to the periodic table?
Here is an example for the electron configuration of chlorine. Reading across the first row, it is $1s^2$. Reading across the second row, it is $2s^22p^6$. Reading across the third row, it is $3s^23p^5$. The electron configuration for chlorine is $1s^22s^22p^63s^23p^5$. 
What is the electron configuration for cadmium?
- Reading across the first row, it is $1s^2$
- Reading across the second row, it is $2s^22p^6$
- Reading across the third row, it is $3s^23p^6$
- Reading across the fourth row, it is $4s^23d^{10}4p^6$
- Reading across the fifth row, it is $5s^24d^{10}$

The electron configuration for cadmium is $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}$
You Try It

Using only a periodic table, write the electron configurations for each of the following elements.

1. Carbon  $1s^22s^22p^2$

2. Calcium  $1s^22s^22p^63s^23p^64s^2$

3. Bromine  $1s^22s^22p^63s^23p^64s^23d^{10}4p^5$
Noble gas notation is a **shorthand** way of writing electron configurations.
Here are some examples of noble gas notations. See if you can figure out how they are done.

Magnesium: \([\text{Ne}]3s^2\)
Bromine: \([\text{Ar}]4s^23d^{10}4p^5\)
Iron: \([\text{Ar}]4s^23d^6\)
Xenon: \([\text{Kr}]5s^24d^{10}5p^6\)
Valence Electrons

The **outermost electrons** play the largest role in determining the **chemical properties** of the elements.

These are the electrons that can be **gained, lost, or shared** in the formation of chemical compounds.

These electrons are known as **valence** electrons.

The valence electrons for the representative elements (elements in groups 1, 2, and 13-18) are the electrons filling the s and p sublevel of the **highest** occupied energy level.

How many valence electrons are in each of the following?

\[ 1s^2 2s^2 2p^5 \quad 7 \]

\[ 1s^2 2s^2 2p^6 3s^1 \quad 1 \]

\[ 1s^2 2s^2 2p^6 3s^2 3p^3 \quad 5 \]
Valence Electrons

Another way to determine the number of valence electrons for the representative elements is by looking at the last digit of the group number or the “A” group number.

Use the periodic table to identify the number of valence electrons in each of the following:

Oxygen  6
Nitrogen  5
Barium  2
Silicon  4
Octet Rule

The octet rule states that elements will **gain**, **lose**, or **share** valence electrons in order to obtain 8 electrons in their s and p orbitals of their outermost energy level.

Noble gas configurations are considered to be **very stable** since their outermost s and p orbitals are full.

Metals tend to **lose** electrons and form **positive ions** (cations). Nonmetals tend to **gain** electrons and form **negative ions** (anions).
Using the Octet Rule and ending electron configurations, predict the charge on the stable ion and the ion symbol of each of the following elements.

<table>
<thead>
<tr>
<th>Element</th>
<th>Noble Gas Notation</th>
<th># of e- gained or lost</th>
<th>Ion Charge</th>
<th>Symbol of Ion</th>
</tr>
</thead>
<tbody>
<tr>
<td>I</td>
<td>[Kr]5s(^2)4d(^{10})5p(^5)</td>
<td>gain 1</td>
<td>-1</td>
<td>I(^-)</td>
</tr>
<tr>
<td>Ca</td>
<td>[Ar]4s(^2)</td>
<td>lose 2</td>
<td>+2</td>
<td>Ca(^{2+})</td>
</tr>
<tr>
<td>P</td>
<td>[Ne]3s(^2)3p(^3)</td>
<td>gain 3</td>
<td>-3</td>
<td>P(^{3-})</td>
</tr>
</tbody>
</table>
Complete the following table.

<table>
<thead>
<tr>
<th>Group Number</th>
<th>Number of Valence Electrons</th>
<th>Charge on Common Ion</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>1</td>
<td>+1</td>
</tr>
<tr>
<td>2</td>
<td>2</td>
<td>+2</td>
</tr>
<tr>
<td>13</td>
<td>3</td>
<td>+3</td>
</tr>
<tr>
<td>14</td>
<td>4</td>
<td>±4</td>
</tr>
<tr>
<td>15</td>
<td>5</td>
<td>-3</td>
</tr>
<tr>
<td>16</td>
<td>6</td>
<td>-2</td>
</tr>
<tr>
<td>17</td>
<td>7</td>
<td>-1</td>
</tr>
<tr>
<td>18</td>
<td>8 (except for helium which has 2)</td>
<td>does not form ions</td>
</tr>
</tbody>
</table>
Isoelectronic

The noble-gas notation for chlorine is \([\text{Ne}]3s^23p^5\). According to the octet rule, chlorine will gain one electron in order to obtain a full outermost energy level.

When chlorine gains one electron, it will become the chloride ion (\(\text{Cl}^-\)). The electron configuration for the chloride ion is \([\text{Ne}]3s^23p^6\). That is the same as the electron configuration for argon. When two species have the same electron configuration they are said to be isoelectronic.

What happens when lithium becomes an ion?

\textbf{Li loses one electron.}

Write an electron configuration for the lithium ion (\(\text{Li}^+\)).

\(1s^2\)

Which neutral element has the same electron configuration?

\textbf{Helium}

Name six ions which are isoelectronic with Neon.

\(\text{N}^3-, \text{O}^2-, \text{F}^-, \text{Na}^+, \text{Mg}^{2+}, \text{Al}^{3+}\)
Periodicity

When the elements are arranged in order of increasing atomic number, there is a periodic recurrence of properties that leads to the grouping of elements in the periodic table.

This periodic recurrence of chemical and physical properties is known as periodicity.

The position of an element in the periodic table can be used to compare periodic trends in atomic radii, electronegativity, ionization energy and ionic radii.
Nuclear Charge and the Shielding Effect

All the periodic trends can be understood in terms of three basic rules.

1. Electrons are attracted to the protons in the nucleus of an atom.
   a. The closer an electron is to the nucleus, the more strongly it is attracted.
   b. The more protons in the nucleus, the more strongly an electron is attracted to the nucleus.

This is known as nuclear charge.

The nuclear charge increases from left to right across a period.
2. Electrons are repelled by other electrons in an atom. So if other electrons are between a valence electron and the nucleus, the valence electron will be less attracted to the nucleus.

The tendency for the electrons in the inner energy levels to block the attraction of the nucleus for the valence electrons is known as the shielding effect.

The shielding effect increases as you go down a group.

The shielding effect remains roughly constant as you go from left to right across a period.
Nuclear Charge and the Shielding Effect

3. Completed p sublevels are very stable. Atoms prefer to add or subtract valence electrons to create completed p sublevels if possible.

Nuclear charge plays an important role in determining period trends.

The shielding effect plays an important role in determining group trends.
Atomic Radii

The atomic radius is half the distance between two nuclei in two adjacent atoms.

The values for the atomic radii of some of the representative elements are given below in picometers.

<p>| | | | | | | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
<th></th>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>H 37</td>
<td>Li 152</td>
<td>Be 112</td>
<td>B 85</td>
<td>C 77</td>
<td>N 75</td>
<td>O 73</td>
</tr>
<tr>
<td>Na 186</td>
<td>Mg 160</td>
<td>Al 143</td>
<td>Si 118</td>
<td>P 110</td>
<td>S 103</td>
<td>Cl 100</td>
</tr>
<tr>
<td>K 227</td>
<td>Cs 197</td>
<td>Ga 135</td>
<td>Ge 122</td>
<td>As 120</td>
<td>Se 119</td>
<td>Br 114</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Generally speaking, how do atomic radii change as you go from left to right across a period? In general the atomic radii decrease.

top to bottom within a group? In general the atomic radii increase.
As effective nuclear charge increases from left to right, outer electrons are held more closely and more strongly to the nucleus.

Top to bottom, outer electrons are held more loosely because they are farther away from the nucleus and the shielding effect increases (electrons in the inner energy levels block the attraction of the nucleus for the valence electrons).
Which atom in each pair would have the larger atomic radii?

a. Li  Cs
b. Li  F
c. K  Br
d. C  Pb
Ionization Energy

The energy required to remove an electron from a gaseous atom is called ionization energy. If the first electron is being removed it is called the first ionization energy.

Once an electron has been removed, the atom becomes a positively charged ion.

The energy required to remove the next electron from the ion is called the second ionization energy and so on.
**Ionization Energy**

The values for the first ionization energy of some of the representative elements are given below in kJ/mol.

<table>
<thead>
<tr>
<th></th>
<th>H</th>
<th>Be</th>
<th>B</th>
<th>C</th>
<th>N</th>
<th>O</th>
<th>F</th>
<th>Ne</th>
</tr>
</thead>
<tbody>
<tr>
<td>1312</td>
<td>Li</td>
<td>520</td>
<td>Be</td>
<td>899</td>
<td>B</td>
<td>801</td>
<td>C</td>
<td>1086</td>
</tr>
<tr>
<td></td>
<td>Na</td>
<td>496</td>
<td>Mg</td>
<td>738</td>
<td>Al</td>
<td>578</td>
<td>Si</td>
<td>786</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>K</td>
<td>419</td>
<td>Cs</td>
<td>590</td>
<td>Ga</td>
<td>579</td>
<td>Ge</td>
<td>762</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Generally speaking, how does first ionization energy change as you go from left to right across a period? **In general it increases.**

In general it decreases.
Explanation

As effective nuclear charge increases from left to right, outer electrons are held more closely and more strongly to the nucleus. This means it will take more energy to remove an electron as you go from left to right across a period.

Top to bottom, outer electrons are held more loosely because they are farther away from the nucleus and the shielding effect increases. This means it will take less energy to remove an electron as you go down the group.
Which atom in each pair would have the larger first ionization energy?

a. Ca  Br
b. Ca  Ba
c. Na  Cs
d. Na  P
Comparing Successive Ionization Energies

The energy required to remove an electron increases as more electrons are removed.

The table below gives the successive ionization energies for the elements in period 3.

<table>
<thead>
<tr>
<th>Element</th>
<th>$I_1$</th>
<th>$I_2$</th>
<th>$I_3$</th>
<th>$I_4$</th>
<th>$I_5$</th>
<th>$I_6$</th>
<th>$I_7$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Na</td>
<td>495</td>
<td>5260</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Mg</td>
<td>735</td>
<td>1445</td>
<td>7330</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Al</td>
<td>580</td>
<td>1815</td>
<td>2740</td>
<td>11600</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Si</td>
<td>780</td>
<td>1575</td>
<td>3220</td>
<td>4350</td>
<td>16100</td>
<td></td>
<td></td>
</tr>
<tr>
<td>P</td>
<td>1060</td>
<td>1890</td>
<td>2905</td>
<td>4950</td>
<td>6270</td>
<td>21200</td>
<td></td>
</tr>
<tr>
<td>S</td>
<td>1005</td>
<td>2260</td>
<td>3375</td>
<td>4565</td>
<td>6950</td>
<td>8490</td>
<td>27000</td>
</tr>
<tr>
<td>Cl</td>
<td>1255</td>
<td>2295</td>
<td>3850</td>
<td>5160</td>
<td>6560</td>
<td>9360</td>
<td>11000</td>
</tr>
<tr>
<td>Ar</td>
<td>1527</td>
<td>2665</td>
<td>3945</td>
<td>5770</td>
<td>7230</td>
<td>8780</td>
<td>12000</td>
</tr>
</tbody>
</table>
## Comparing Successive Ionization Energies

<table>
<thead>
<tr>
<th>Element</th>
<th>( \text{I}_1 ) (kJ/mol)</th>
<th>( \text{I}_2 ) (kJ/mol)</th>
<th>( \text{I}_3 ) (kJ/mol)</th>
<th>( \text{I}_4 ) (kJ/mol)</th>
<th>( \text{I}_5 ) (kJ/mol)</th>
<th>( \text{I}_6 ) (kJ/mol)</th>
<th>( \text{I}_7 ) (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Na</td>
<td>495</td>
<td>5260</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Mg</td>
<td>735</td>
<td>1445</td>
<td>7330</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Al</td>
<td>580</td>
<td>1815</td>
<td>2740</td>
<td>11600</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Si</td>
<td>780</td>
<td>1575</td>
<td>3220</td>
<td>4350</td>
<td>16100</td>
<td></td>
<td></td>
</tr>
<tr>
<td>P</td>
<td>1060</td>
<td>1890</td>
<td>2905</td>
<td>4950</td>
<td>6270</td>
<td>21200</td>
<td></td>
</tr>
<tr>
<td>S</td>
<td>1005</td>
<td>2260</td>
<td>3375</td>
<td>4565</td>
<td>6950</td>
<td>8490</td>
<td>27000</td>
</tr>
<tr>
<td>Cl</td>
<td>1255</td>
<td>2295</td>
<td>3850</td>
<td>5160</td>
<td>6560</td>
<td>9360</td>
<td>11000</td>
</tr>
<tr>
<td>Ar</td>
<td>1527</td>
<td>2665</td>
<td>3945</td>
<td>5770</td>
<td>7230</td>
<td>8780</td>
<td>12000</td>
</tr>
</tbody>
</table>

The second ionization energy for sodium is much higher than the first ionization energy for sodium. Why do you think this is so?

After the first electron has been removed, the sodium ion formed has an electron configuration that is isoelectronic with that of a noble gas and is therefore very stable.
 Ionic Radii

The *ionic radius* is the radius of a *cation* or *anion*.

When the atom loses or gains electrons, the resulting ion *changes* in size from the original atom.

Metals tend to *lose* electrons and form *cations*.

Nonmetals tend to *gain* electrons and form *anions*.

The names of monatomic *anions* end in *–ide*. Examples: sulfide, phosphide, fluoride
Generally speaking, how does the size of the cation change as you go from left to right across a period? In general it decreases.

top to bottom within a group? In general it increases.
How does the size of the cation compare to the size of the neutral atom?

The cation is always smaller than the neutral atom.
Cations are always smaller than their parent atoms because the electrons lost upon formation of a cation vacate the outermost orbitals, decreasing the size of the ion. Additionally there are fewer electron-electron repulsions. The effective nuclear charge increases.

Example: Sodium
Sodium Atom: $1s^22s^22p^63s^1$
Sodium Ion: $1s^22s^22p^6$
Generally speaking, how does the size of the anion change as you go from left to right across a period? **In general it decreases.**

How does the size of the anion change as you go from top to bottom within a group? **In general it increases.**
How does the size of the anion compare to the size of the neutral atom?
The anion is always larger than the neutral atom.
Explanation

Anions are always larger than their parent atoms because additional electrons cause increased electron-electron repulsions causing the electrons spread out more in space. The effective nuclear charge decreases.

Example: Chlorine
Chlorine atom: \(1s^22s^22p^63s^23p^5\)
Chloride ion (Cl\(^-\)): \(1s^22s^22p^63s^23p^6\)
Comparing Ions

Which ion would you expect to have the smaller ionic radius: Li$^+$ or Be$^{2+}$? Why?

The lithium atom contains 3 p$^+$ and 3e$^-$. The lithium atom loses 1 e$^-$ to form the lithium ion (Li$^+$) (3p$^+$ and 2e$^-$).

The beryllium atom contains 4p$^+$ and 4e$^-$. The beryllium atom loses 2 e$^-$ to form the beryllium ion (Be$^{2+}$) (4 p$^+$ and 2e$^-$)

The nuclear charge is greater for the beryllium ion, so it would be smaller than the lithium ion. The electron-electron repulsion is also less.
Comparing Ions

Which ion would you expect to have the smaller ionic radius: $\text{F}^-\text{ or } \text{O}^{2-}$? Why?

The fluorine atom contains $9\text{p}^+\text{ and } 9\text{e}^-$. The fluorine atom gains $1\text{ e}^-$ to form the fluoride ion ($\text{F}^-$) ($9\text{p}^+\text{ and } 10\text{e}^-$).

The oxygen atom contains $8\text{p}^+\text{ and } 8\text{e}^-$. The oxygen atom gains $2\text{ e}^-$ to form the oxide ion ($\text{O}^{2-}$) ($8\text{ p}^+\text{ and } 10\text{e}^-$).

The nuclear charge is greater for the fluoride ion, so it would be smaller than the oxide ion. The electron-electron repulsion is also less for the fluoride ion.
Which particle in each pair would be larger?

a. Lithium atom  Lithium ion
b. Fluorine atom  Fluoride ion

c. Sodium ion  Magnesium ion
d. Sulfide ion  Oxide ion
Electronegativity

Electronegativity refers to the tendency for an atom to attract electrons to itself when it is chemically combined with another element.

The electronegativity values for some of the representative elements are given below.

<table>
<thead>
<tr>
<th></th>
<th>H 2.1</th>
<th>Be 1.5</th>
<th>B 2.0</th>
<th>C 2.5</th>
<th>N 3.0</th>
<th>O 3.5</th>
<th>F 4.0</th>
<th>Ne -</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li</td>
<td>1.0</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Na</td>
<td>0.9</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>K</td>
<td>0.8</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Ca</td>
<td>1.0</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Ga</td>
<td>1.6</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Ge</td>
<td>1.8</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>As</td>
<td>2.0</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Se</td>
<td>2.4</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Br</td>
<td>2.8</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Kr</td>
<td>3.0</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Generally speaking, how does electronegativity change as you go from left to right across a period? Increases

top to bottom within a group? Decreases
a. The noble gases do not generally have electronegativity values. Why do you think this is so?

   The noble gases do not generally form compounds.

b. Which group of elements has the highest electronegativity values?

   The halogens have the highest electronegativity value.

c. Which group of elements has the lowest electronegativity values?

   The alkali metals have the lowest electronegativity value.
Summary of Periodic Trends

- Shielding is constant
- Nuclear charge increases
- Atomic radius decreases
- Ionization energy increases
- Electron affinity increases
- Electronegativity increases

Ionic size (cations) decreases

Ionic size (anions) decreases