The concept of the atom dates back to ancient Greece around 400 B.C. Democritus and Aristotle founded the “particle theory of matter” Atoma – “indivisible”
THE LAWS (circa 1800)

- Law of Conservation of Mass: Mass is neither created nor destroyed during chemical reactions.
- Law of Definite Proportions: A chemical compound contains the same proportions of elements by mass regardless of sample size or source location.
  - Water (H₂O) is water (H₂O) everywhere!
THE LAWS (circa 1800)

- Law of Multiple Proportions: If two or more compounds are made of the same two elements, then the ratio of masses of the 2nd element combined with a certain mass of the 1st element is always a ratio of small whole numbers.

- What does this mean!?...
Law of Multiple Proportions… for real
Compounds made from the same elements have different ratios and formulas
- $\text{H}_2\text{O}$ vs. $\text{H}_2\text{O}_2$
- $\text{NO}$ vs. $\text{NO}_2$
- $\text{CH}_4$ vs. $\text{C}_3\text{H}_8$

Summed up in 5 points:
Dalton’s Atomic Theory

My theory states:

1. All elements are made up of tiny particles called atoms.
2. Atoms of a given element are alike.
3. Atoms of different elements are different.
4. Chemical changes take place when atoms link up with or separate from one another.
5. Atoms are not created or destroyed by chemical change.

Dalton’s Model
- hydrogen atom
- oxygen atom
Much of Dalton’s Theory stands presently imperfect because of technological limitations of the time:
- No knowledge of isotopes, ions, radioactivity, subatomic particles
MODERN ATOMIC THEORY and ATOMIC STRUCTURE

- The 19th century advance atomic theory with the discovery of the electron in 1897
- Cathode Ray Tubes (CRTs) demonstrated the existence of a charged particle in all elements
CATHODE RAYS

- An electric charge through a low pressure gas causes the surface opposite the cathode to glow.
- The ray of “light” traveled from cathode (-) to anode (+) and can be pushed” by magnets.
The conclusion is that cathode rays are composed of negatively charged particles.

All metals demonstrated the same phenomenon.

The identical negative charged particles are the ELECTRON.
THOMSON’S MODEL

- The model of the atom was redesigned by 1900.
- The “Plum Pudding” model added small negative particles to the surface of a solid atomic (positive) sphere.
RUTHERFORD’S MODEL

- Rutherford designs the “Gold Foil Experiment” and discovers the small positively charged nucleus.
- Most of the atom is “empty”
- The nucleus is extremely small and dense and positive
THE NUCLEUS

- The nucleus contains PROTONS and NEUTRONS
- NEUTRONS are more massive than protons
- NEUTRONS have no electric charge and stabilize the repulsive forces in the nucleus
- NUCLEAR FORCE: P-P, N-N, P-N forces that keep the nucleus together
COUNTING ATOMS

- **ATOMIC NUMBER** of an element is the number of protons in each atom of the element.
- Each element **ATOMIC NUMBER** is a unique whole number (1-118).
ISOTOPES

- ISOTOPES are atoms of the same element that have different masses
- ISOTOPES are identified by MASS NUMBER
- MASS NUMBER is the total number of particles in the nucleus (Protons & Neutrons)

Three Isotopes of Hydrogen

- Protium: $^1\text{H}$
- Deuterium: $^2\text{H}$
- Tritium: $^3\text{H}$

Carbon
- 6 Protons
- Mass Number: 6
- 6 Neutrons

Carbon-13
- 6 Protons
- Nuclear number: 6
- 7 Neutrons
- Mass Number: 13

Carbon-14
- 6 Protons
- Nuclear number: 6
- 8 Neutrons
- Mass Number: 14
ISOTOPE NOTATION

- **HYPHEN NOTATION** uses the element name or symbol, a hyphen, and the mass number.

- **NUCLEAR NOTATION** uses the element symbol and superscripts and subscripts to indicate mass number and atomic number.

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Mass Number</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>1</td>
<td>hydrogen-1 (protium)</td>
</tr>
<tr>
<td></td>
<td>2</td>
<td>hydrogen-2 (deuterium)</td>
</tr>
<tr>
<td></td>
<td>3</td>
<td>hydrogen-3 (tritium)</td>
</tr>
</tbody>
</table>
RELATIVE ATOMIC MASS

- RELATIVE ATOMIC MASS is based on the carbon-12 atom.
- The unit of measure is the atomic mass unit (amu).
- $1 \text{ amu} = 1/12$ the mass of the carbon-12 atom.
- $1 \text{ amu} = 1.66 \times 10^{-24} \text{ g} = (6.02 \times 10^{23})^{-1}$
- The very small amu makes the individual masses of atoms appear without exponential notation.
RELATIVE ATOMIC MASS

EXAMPLE:

- Oxygen-16 has a mass of $2.656 \times 10^{-23}$ g
- AWKWARD AND TOO SMALL
- Proton and neutron masses are about, but not equal to 1 amu
- This is why mass numbers and relative atomic masses are close, but not equal
- Oxygen-16: $g \rightarrow \text{amu}$
- $(2.656 \times 10^{-23} \text{ g})/ (1.66 \times 10^{-24} \text{ amu/g}) = 16 \text{ amu}$
  (actual = 15.995 amu)
The weighted average of the atomic masses of the naturally occurring isotopes of an element is the AVERAGE ATOMIC MASS.
MASS & ATOMS

- THE MOLE is equal to Avogadro’s Number
- \(6.02 \times 10^{23} = 1 \text{ mole “things”}\)
- Based on carbon-12 atom
- 12 g of carbon-12 = \(6.02 \times 10^{23}\) atoms = 1 mole
- MOLAR MASS is the mass on 1 mole of a pure substance (element or compound)
- MOLAR MASS units are (g/mol) and is equal to the average atomic mass value
Molar Mass

- The mass of an atom in amu is also equal to the mass in grams of one mole of the element.

<table>
<thead>
<tr>
<th>Element</th>
<th>Atomic mass</th>
<th>Molar mass</th>
<th>Number of atoms</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>1.01 amu</td>
<td>1.01 g/mol</td>
<td>$6.022 \times 10^{23}$</td>
</tr>
<tr>
<td>Mg</td>
<td>24.31 amu</td>
<td>24.31 g/mol</td>
<td>$6.022 \times 10^{23}$</td>
</tr>
<tr>
<td>Na</td>
<td>22.99 amu</td>
<td>22.99 g/mol</td>
<td>$6.022 \times 10^{23}$</td>
</tr>
</tbody>
</table>

Molecular Weight of CH₄

\[
\text{CH}_4 : \quad \begin{align*}
\text{C} & : 12.01 \text{ g/mol} \\
\text{H} & : (1.01 \text{ g/mol} \times 4) = 4.04 \text{ g/mol}
\end{align*}
\]

\[
= 12.01 \text{ g/mol} + 4.04 \text{ g/mol} = 16.05 \text{ g/mol}
\]
PRACTICE!

YOU'RE A 10?

ON THE PH SCALE MAYBE, CAUSE YOU BASIC
THE MODERN MODEL

- Rutherford's model failed to explain how electrons didn't collapse toward the nucleus.
- Advances in technology and physics helped to solve the mystery by the end of the 1920's.
LIGHT & THE ELECTRON

- LIGHT was considered only as a WAVE prior to 1900
- The 20th century revealed that LIGHT exhibits both WAVE and PARTICLE behavior (DUALITY)
- A new understanding of LIGHT solved the mystery of the never collapsing electrons in atoms
LIGHT DETAILS

- **LIGHT** is a general term for a type of **ELECTROMAGNETIC RADIATION (EMR)**
- EMR exhibits wavelike behavior as it travels through space (matter & vacuum)
- The **ELECTROMAGNETIC SPECTRUM (EMS)** includes all forms of EMR
WAVE PROPERTIES

- WAVELENGTH is the length of a wave, measured by the distance between corresponding points on adjacent waves.
- Measured in (m) or (nm)

- FREQUENCY is the number of waves that pass a given point in a specific time.
- Measured in (Hz) or s\(^{-1}\)
WAVE PROPERTIES

- The SPEED OF LIGHT through a medium or vacuum is constant
- WAVELENGTH and FREQUENCY are inversely proportional (speed = wavelength \* frequency)
- The SPEED OF LIGHT in a vacuum is 3.00 x 10^8 m/s
The PHOTOELECTRIC EFFECT is the emission (loss) of electrons from a metal when light of a certain frequency shines on a metal. Light particles with enough energy can knock an electron from an atom.
LIGHT AS A PARTICLE

- The frequency of the light is important for the photoelectric effect.
- The emitted electrons represent a QUANTUM of energy (minimum quantity of energy that can be gained or lost by an atom).
- QUANTA are finite and not continuous like a wave.
- Max Planck discovered the relationship between energy and frequency \( (E = hv) \).
WHAT ABOUT THE ELECTRON?

- The short… ELECTRONS are like LIGHT
- ELECTRONS exhibit both WAVE & PARTICLE behavior
- ELECTRONS as a WAVE solve the mystery of the “never collapsing electron”
BOHR & HYDROGEN

- The potential energy of atoms in a low pressure gas tube increases when electricity flow through.
- GROUND STATE is the lowest energy state of an atom.
- EXCITED STATE occurs when atoms have a potential energy higher than the ground state.
BOHR & HYDROGEN

- The “pink” light emitted by hydrogen can be examined with a spectrometer to produce a LINE-EMISSION SPECTRUM.
- The spectrum is not continuous as it would be predicted by classical theory.
- QUANTUM THEORY attempts to explain the line spectrum observed.
DeBroglie (1924) proposed the electron as a wave.

New experiments demonstrated electrons can be “bent” or “diffracted” and show “interference” as only waves do.
ELECTRON WAVES!

- Heisenberg (1927) addresses the ability to detect electrons using photons
- Photons have about the same energy as an electron and “knock” electrons “off course” to detect them
- HEISENBERG UNCERTAINTY PRINCIPLE: it is impossible to simultaneously know both the position and velocity of an electron
Hmmm...
Schroedinger (1926) developed an equation that treats electrons as waves.

Electron energy is the “output” of Schrodinger’s Equation.

Schrodinger and Heisenberg laid the foundation for QUANTUM THEORY.

QUANTUM THEORY describes the wave properties of electrons mathematically.
A careful analysis of the process of observation in atomic physics has shown that the subatomic particles have no meaning as isolated entities, but can only be understood as interconnections between the preparation of an experiment and the subsequent measurement.

Erwin Schrödinger
Hydrogen Wave Function

\[ \psi_{nlm}(r, \theta, \varphi) = \left( \frac{2}{m \omega} \right)^{\frac{1}{2}} \frac{(n-l-\frac{1}{2})!}{2(n+l+1)!} \frac{\gamma}{\pi} \frac{\gamma}{\pi} \frac{(n+l+1)!}{n!} \frac{1}{\sqrt{\lambda}} \rho^{l+\frac{1}{2}} \cdot Y_{lm}(\theta, \varphi) \]
ORBITALS!

- QUANTUM NUMBERS specify the properties of atomic orbitals and the electrons in the orbitals.
- ORBITALS are 3D regions of space around the nucleus that indicate the probable location of electrons.
ORBITALS!

s-block
1
1s
2s
3s
4s
5s
6s
7s

d-block
3
4
5
6
7
8
9
10
11
12

p-block
13
14
15
16
17
18
2p
3p
4p
5p
6p
7p

f-block
4f
5f
HONORS CHEMISTRY
Unit B: Atoms, Electrons, and The Periodic Table
CHAPTER FIVE: PERIODIC LAW
The Periodic Table begins to take form around 1860 with 60 known elements after the First International Congress of Chemists. Elements were initially organized by atomic mass and properties (Cannizzaro 1860). Mendeleev is the first to organize by periodic pattern. The modern Periodic Table is crafted by Mendeleev in 1869.
Mendeleev left gaps for elements yet to be discovered and was proved right by 1886.

Mendeleev organized his table by increasing atomic mass – MISTAKE - eventually corrected by Moseley in 1911.
PERIODIC TABLE HISTORY

- Moseley recognizes the Periodic Law is a function of Atomic Number, not Mass.
- PERIODIC LAW states that chemical and physical properties of elements are functions of their atomic number.
Modern Periodic Table of the Elements arranges the elements in order of increasing atomic number so that elements with similar properties fall into groups.

![Periodic Table Image]

### Modern Periodic Tables

<table>
<thead>
<tr>
<th>Group</th>
<th>Elements</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>H, He</td>
</tr>
<tr>
<td>2</td>
<td>Li, Be, B, C, N, O, F, Ne</td>
</tr>
<tr>
<td>3</td>
<td>Na, Mg, Al, Si, P, S, Cl, Ar</td>
</tr>
<tr>
<td>4</td>
<td>K, Ca, Sc, Ti, V, Cr, Mn, Fe, Co, Ni, Cu, Zn, Ga, Ge, As, Se, Br, Kr</td>
</tr>
<tr>
<td>5</td>
<td>Rb, Sr, Y, Zr, Nb, Mo, Tc, Ru, Rh, Pd, Ag, Cd, In, Sn, Sb, Te, I, Xe</td>
</tr>
<tr>
<td>6</td>
<td>Cs, Ba, La, Ce, Pr, Nd, Pm, Sm, Eu, Gd, Tb, Dy, Ho, Er, Tm, Yb, Lu</td>
</tr>
<tr>
<td>7</td>
<td>Fr, Ra, Ac, Th, Pa, U, Np, Pu, Am, Cm, Bk, Cf, Es, Ac, Th, Pa, U</td>
</tr>
</tbody>
</table>

- *Transition Metals*
Electron Configurations and The Periodic Table

- The Periodic Table is organized into BLOCKS
- S-Block – Groups 1-2 (Alkali & Alkaline Earth Metals)
- P-Block – Groups 13-18 (Halogens & Noble Gases)
- D-Block – Groups 3-12 (Transition Metals)
- F-Block – Lanthanides & Actinides
ORBITALS!

s-block

- 1s
- 2s
- 3s
- 4s
- 5s
- 6s
- 7s

d-block

- 3d
- 4d
- 5d
- 6d

p-block

- 2p
- 3p
- 4p
- 5p
- 6p
- 7p

f-block

- 4f
- 5f
### Periodic Table of Elements

#### Groups

<table>
<thead>
<tr>
<th>Period</th>
<th>IA</th>
<th>IIA</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>H</td>
<td>He</td>
</tr>
<tr>
<td>2</td>
<td>Li</td>
<td>Be</td>
</tr>
<tr>
<td>3</td>
<td>Na</td>
<td>Mg</td>
</tr>
<tr>
<td>4</td>
<td>K</td>
<td>Ca</td>
</tr>
<tr>
<td>5</td>
<td>Rb</td>
<td>Sr</td>
</tr>
<tr>
<td>6</td>
<td>Cs</td>
<td>Ba</td>
</tr>
<tr>
<td>7</td>
<td>Fr</td>
<td>Ra</td>
</tr>
</tbody>
</table>

#### Lanthanide Series

<table>
<thead>
<tr>
<th>LA</th>
<th>Ce</th>
<th>Pr</th>
<th>Nd</th>
<th>Sm</th>
<th>Eu</th>
<th>Gd</th>
<th>Tb</th>
<th>Dy</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>140.1</td>
<td>140.9</td>
<td>144.2</td>
<td>150.4</td>
<td>152.0</td>
<td>157.2</td>
<td>158.9</td>
<td>162.5</td>
</tr>
</tbody>
</table>

#### Actinide Series

<table>
<thead>
<tr>
<th>Act</th>
<th>Ac</th>
<th>Th</th>
<th>Pa</th>
<th>U</th>
<th>Pu</th>
<th>Am</th>
<th>Cm</th>
<th>Bk</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>227</td>
<td>232</td>
<td>231</td>
<td>238</td>
<td>244</td>
<td>243</td>
<td>247</td>
<td>247</td>
</tr>
</tbody>
</table>

#### Periodic Table Key

- Other Non-Metals
- Alkali Metals
- Alkalai Earth Metals
- Transition Metals
- Other Metals
- Metalloids
- Halogens
- Noble Gases
PERIODIC TRENDS

- TRENDS are patterns observed in periods and groups of the Periodic Table.
- Typically described as LEFT to RIGHT and/or TOP to BOTTOM and as either INCREASING or DECREASING.
- SIZE TRENDS (2): Atomic Radii & Ionic Radii.
Periodic Trends
ATOMIC RADII

- ATOMIC RADII is half the distance between the nuclei of identical bonded atoms.
- Measures from the nucleus to the approximate edge of an atom.
- Very small distances usually given in picometers (pm) \( \rightarrow 1 \text{ pm} = 10^{-12} \text{ m} \).
ATOMIC RADII

- PERIOD TREND: atoms tend to decrease in size across a period (L → R)
- WHY?... Greater nuclear charge (+) without additional energy levels added
- GROUP TREND: atoms tend to increase in size down a group (T → B)
- WHY?... Additional energy levels added
IONS are charged atoms (either + or -)
IONS usually form when an atom gains or loses electrons to achieve a greater stability (octet)
CATIONS are positively charged ions (+, 2+, 3+, ...)
ANIONS are negatively charged ions (-, 2-, 3-)
ION sizes change due to the loss or gain of electrons and nuclear charge interaction
IONIC RADII is measured in (pm)
IONIC RADII

- **PERIOD TREND:** atoms tend to decrease in size across a period (L → R) (same as atomic radii)
- **GROUP TREND:** atoms tend to increase in size down a group (T → B) (same as atomic radii)
- **ATOM vs. ION SIZE CHANGE:**
  - **CATIONS:** for the same element cations decrease in size
  - **ANIONS:** for the same element anions increase in size
VALENCE ELECTRONS

- VALENCE ELECTRONS are the electrons that are gained, lost or shared in the formation of chemical compounds.
- These are the S & P orbital electrons of the highest energy level.
- Octets (8) are the max and ideal.
- Located at the “edge” of atoms and most influenced by other atoms.
VALENCE ELECTRONS

Number of Valence Electrons
VALENCE ELECTRONS

PERIODIC TABLE ELEMENTS 1–20

<table>
<thead>
<tr>
<th>Hydrogen</th>
<th>Helium</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>He</td>
</tr>
<tr>
<td>Lithium</td>
<td>Beryllium</td>
</tr>
<tr>
<td>Li</td>
<td>Be</td>
</tr>
<tr>
<td>Sodium</td>
<td>Magnesium</td>
</tr>
<tr>
<td>Na</td>
<td>Mg</td>
</tr>
<tr>
<td>Potassium</td>
<td>Calcium</td>
</tr>
<tr>
<td>K</td>
<td>Ca</td>
</tr>
</tbody>
</table>

Group 1
Period 1
(H in filling)

Group 14
Period 2
(2n in filling)

Group 17
Period 3
(3n in filling)

Group 18

He, Ne
IONIZATION ENERGY

- IONIZATION ENERGY is the energy required to remove an electron from a neutral atom.
- $\text{IE}_1$ (1st electron), $\text{IE}_2$ (2nd electron), $\text{IE}_3$ (3rd electron)
- The energy measurements are made on isolated atoms in the gas phase.
- Measured in kJ/mol.
- $\text{A} + \text{energy} \rightarrow \text{A}^+ + \text{e}^-$
- Energy is required because of the nucleus (+) & electron (-) attraction within the atom.
<table>
<thead>
<tr>
<th>Ionization Energy</th>
</tr>
</thead>
<tbody>
<tr>
<td>Increasing</td>
</tr>
<tr>
<td>Decreasing</td>
</tr>
</tbody>
</table>
## IONIZATION ENERGY

### TABLE 6.2 Higher Ionization Energies (kJ/mol) for Third-Row Elements

<table>
<thead>
<tr>
<th>( E_1 ) Number</th>
<th>Na</th>
<th>Mg</th>
<th>Al</th>
<th>Si</th>
<th>P</th>
<th>S</th>
<th>Cl</th>
<th>Ar</th>
</tr>
</thead>
<tbody>
<tr>
<td>( E_{11} )</td>
<td>496</td>
<td>738</td>
<td>578</td>
<td>787</td>
<td>1,012</td>
<td>1,000</td>
<td>1,251</td>
<td>1,520</td>
</tr>
<tr>
<td>( E_{12} )</td>
<td>4,562</td>
<td>1,451</td>
<td>1,817</td>
<td>1,577</td>
<td>1,903</td>
<td>2,251</td>
<td>2,297</td>
<td>2,665</td>
</tr>
<tr>
<td>( E_{13} )</td>
<td>6,912</td>
<td>7,733</td>
<td>2,745</td>
<td>3,231</td>
<td>2,912</td>
<td>3,361</td>
<td>3,822</td>
<td>3,931</td>
</tr>
<tr>
<td>( E_{14} )</td>
<td>9,543</td>
<td>10,540</td>
<td>11,575</td>
<td>4,356</td>
<td>4,956</td>
<td>4,564</td>
<td>5,158</td>
<td>5,770</td>
</tr>
<tr>
<td>( E_{15} )</td>
<td>13,353</td>
<td>13,630</td>
<td>14,830</td>
<td>16,091</td>
<td>6,273</td>
<td>7,013</td>
<td>6,540</td>
<td>7,238</td>
</tr>
<tr>
<td>( E_{16} )</td>
<td>16,610</td>
<td>17,995</td>
<td>18,376</td>
<td>19,784</td>
<td>22,233</td>
<td>8,495</td>
<td>9,458</td>
<td>8,781</td>
</tr>
<tr>
<td>( E_{17} )</td>
<td>20,114</td>
<td>21,703</td>
<td>23,293</td>
<td>23,783</td>
<td>25,397</td>
<td>27,106</td>
<td>11,020</td>
<td>11,995</td>
</tr>
</tbody>
</table>
IONIZATION ENERGY

- PERIOD TREND: energies tend to increase in size across a period (L → R)
- GROUP TREND: energies tend to decrease in size down a group (T → B)
- There are exceptions!
- Successive energies (IE$_1$ vs. IE$_2$ vs. IE$_3$) tend to increase and will become exceptionally large after an octet has been disturbed
ELECTRON AFFINITY

- ELECTRON AFFINITY is the energy change that occurs when an electron is acquired by a neutral atom.
- Most atoms release energy when an electron is acquired (measured in kJ/mol).
- Some atoms must be “forced” to gain an electron with the addition of energy (unstable and recorded as “zero”).
- $A + e^- \rightarrow A^- + \text{energy} (-\#)$ OR $A + e^- + \text{energy} \rightarrow A^- (0)$
ELECTRON AFFINITY

Magnitude of electron affinity (kJ/mol), s-, p-, and d-block elements

Increasing

Electron Affinity (kJ/mol)
ELECTRON AFFINITY

- **PERIOD TREND:** affinity tends to increase across a period (L $\rightarrow$ R) (peaks at Group 17)
- **GROUP TREND:** affinity values are fairly irregular by group (T $\rightarrow$ B)
- There are exceptions everywhere!
- The “most negative values” signify the highest affinity to acquire an electron (-100 to -400)
- Metals tend to have “low” affinities (0 to 100)
ELECTRONEGATIVITY

- ELECTRONEGATIVITY is a measure of the ability of an atom in a compound to attract electrons from other atoms in the compound.
- Results in “uneven sharing” of electrons.
- Fluorine is the most electronegative element.
- Based on a 4.0 scale (F = 4.0).
### Pauling Electronegativity Values

<table>
<thead>
<tr>
<th>Element</th>
<th>Electronegativity</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>2.1</td>
</tr>
<tr>
<td>Li</td>
<td>1.0</td>
</tr>
<tr>
<td>Be</td>
<td>1.6</td>
</tr>
<tr>
<td>Na</td>
<td>0.9</td>
</tr>
<tr>
<td>Mg</td>
<td>1.3</td>
</tr>
<tr>
<td>K</td>
<td>0.8</td>
</tr>
<tr>
<td>Ca</td>
<td>1.0</td>
</tr>
<tr>
<td>Sc</td>
<td>1.4</td>
</tr>
<tr>
<td>Ti</td>
<td>1.5</td>
</tr>
<tr>
<td>V</td>
<td>1.6</td>
</tr>
<tr>
<td>Cr</td>
<td>1.7</td>
</tr>
<tr>
<td>Mn</td>
<td>1.5</td>
</tr>
<tr>
<td>Fe</td>
<td>1.8</td>
</tr>
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Al  1.5  Si  1.9  P  2.2  S  2.6  Cl  3.0  Ar
K  0.8  Ca  1.0  Sc  1.4  Ti  1.5  V  1.6  Cr  1.7  Mn  1.5  Fe  1.8  Co  1.9  Ni  1.9  Cu  1.9  Zn  1.6  Ga  1.8  Ge  2.0  As  2.2  Se  2.6  Br  2.8  Kr
Rb 0.8  Sr  0.9  Y  1.2  Zr  1.3  Nb  1.6  Mo  2.2  Tc  1.9  Ru  2.2  Rh  2.3  Pd  2.2  Ag  1.9  Cd  1.7  In  1.8  Sn  2.0  Sb  2.1  Te  2.1  I  2.5  Xe

Charles E. Sundin, University of Wisconsin-Platteville
ELECTRONEGATIVITY

Dipole moment

Electronegativity
ELECTRONEGATIVITY

- PERIOD TEND: increase across a period (L → R)
- GROUP TEND: decrease down a group (T → B)
- Fluorine is the most electronegative element
- Based on a 4.0 scale (F = 4.0)
- Nonmetals have EN values higher than metals