The following is covered in this Unit:

A. History and some Laws *(not on the test)*
B. Subatomic particles, Atomic number, atomic mass, and Isotopes
C. Energy levels and electron configurations
D. The modern periodic table and periodic trends
E. Writing Lewis structures of ionic and covalent compounds
F. London forces and Hydrogen bonds

**A. The History: Atoms and the Periodic Table**

*BBC has a great documentary on the history of the elements – it is well worth watching (the link is on the website). You will not be tested on the history, but still read it (or watch the documentary)- also “the history of atomic chemistry” by Crash Course Chemistry#37*

**Early Ideas**

*~470 BC Empedocles:*

- Early belief was that all matter was composed of 4 elements: earth, air, fire and water

*~440 BC Democritus*

- Democritus came up with the idea that all matter was made up of tiny *indivisible* particles which he called atoms (*Greek for indivisible*)
- He thought that these particles were in constant motion and combined with each other in various ways
- Not based on scientific observations

*384-270 BC - Plato and Aristotle:*

- Aristotle – opposed this theory and supported early belief that all matter was composed of earth, air, fire and water.
- Aristotle’s influence meant that this theory was the one supported until the 1600’s AD

*1500’s Galileo* believed that a new substance thru chemical change involved the re-arrangement of parts too small to be seen

*17*th *century – Robert Boyle and Isaac Newton* – used atomic concepts to interpret physical phenomenon

*Late 1700’s, early 1800’s -Dalton’s model of the atom*
In the late 1700’s, early 1800’s an English schoolmaster – John Dalton renewed the idea of atoms and proposed a model based on facts and experimental evidence – His idea was based on a different kind of atom for each element

Dalton’s atomic model:
1) Elements are made of tiny, indivisible particles called atoms
2) Atoms of the same element are alike in mass and size
3) Atoms of different elements have different masses and sizes
4) Chemical compounds formed by union of 2 or more atoms of different elements
5) Atoms combine to form compounds in simple ratios
6) Atoms of 2 elements combine in different ratios to form more than 1 compound

This model was a landmark in chemistry – the major premises of Dalton’s theory are still valid with some modifications
1) Atoms are composed of subatomic particles
2) Not all atoms of 1 element have the same mass (i.e. isotopes)
3) Atoms can be decomposed

Some Laws...

A compound will always contain the same elements in the same proportion by mass
i.e., $\text{H}_2\text{O} = 11.2\% \text{ H} \text{ and } 88.8\% \text{ O}$ by mass

Law of definite composition states: a compound always contains 2 or more elements combined in a definite proportion by mass. Composition of a compound will always be the same – no matter how it is formed

Law of Multiple proportions states: atoms of 2 or more elements may combine in different ratios to produce more than one compound.

Different compounds will be unique.

For example combining hydrogen and oxygen in different ratios makes different compounds
- 2 H and 1 O makes $\text{H}_2\text{O} = \text{water}$
- 2 H and 2 O makes $\text{H}_2\text{O}_2 = \text{hydrogen peroxide}$

some examples
- Ethyl alcohol = $\text{C}_2\text{H}_6\text{O}$
- Glucose = $\text{C}_6\text{H}_{12}\text{O}_6$
Discovery of ions and properties of electric charge...

- English Scientist in the early 1800’s Michael Faraday discovered that some substances dissolved in water conducted an electric current
  - he also noticed that certain compounds decomposed into their elements when you pass an electric current thru them – atoms of some were attracted to the positive and others to the negative
  - Therefore - he concluded that atoms can be electrically charged –
  - Called ‘ions’ = (Greek for wanderer)
- Any moving charge is an electric current → moving charge must travel thru a substance known as a conducting medium – i.e. wire
- Swedish scientist Arrhenius in the late 1800’s – concluded that an ion was an atom or group of atoms carrying a positive or negative charge
  Thru experiment he came up with
  - Cations = positively charged ion that moves towards a negative pole (cathode)
  - Anion = negatively charge ions that move towards a positive pole (anode)
- Then an Irish named Stoney used Faraday’s and Arrhenius’s ideas in the late 1800’s and concluded that unit of electricity must be associated with atoms – he called it the electron.
- In 1891 he had no way to prove it so it wasn’t until 1897 that JJ Thompson was able to experimentally show the existence of the electron.

The properties of electric charge:

- Charges are 2 types: positive or negative
- Unlike charges attract, like charges repel
- Charge can be transferred between objects by contact or induction (Electrostatic induction is a method to create or generate static electricity in a material by bringing an electrically charged object near it. This causes the electrical charges to be redistributed in the material, resulting in one side having an excess of either positive (+) or negative (−) charges.)
- Less distance between 2 charges = greater force of attraction between unlike and repulsion between like forces

History of Subatomic particles...

a. How the subatomic particles were discovered:

The development of atomic theory advanced mainly because of the invention of new instruments:
• Crookes tube – developed by Sir William Crookes – was able to discover info about the subatomic structure of the atom
• The Crookes tube emits cathode rays
• JJ Thomson demonstrated that cathode rays:
  1. Travel in straight lines
  2. Are negative in charge
  3. Are deflected by electric and magnetic fields
  4. Produce shadows
  5. Can move paddle wheel

  This experiment led to the discovery of the fundamental unit of charge = the electron

b. Thomson model of the atom:

*Thomson found that:*

• Electron = has a negative charge – has a relative charge of -1 (its real charge is too cumbersome therefore given a relative charge of -1) also has a relative mass of 0.0005 AMU (atomic mass unit) but so small considered to be 0.
• Proton – Thompson discovered the nature of the proton – has a relative charge of +1 and a relative mass of 1 AMU

*His model:*
Electrons are negatively charged particles embedded in positively charged atomic sphere – a neutral atom could become charged by gaining or losing electrons

For example: positive ions meant that a neutral atom had lost electrons, therefore, an atom with a net charge of +1 had lost one electron, +2 had lost 2 etc..

  Whereas an atom with a net charge of -1 had gained 1 electrons, -2 gained 2 electrons etc...

c. Bohr and Planck’s further refinements to the idea of the atom:

*First a bit about electromagnetic waves – what Bohr and others use to help explain their ideas:*
Various ideas over time have been proposed to explain how energy is transferred from place to place -like sound travel, heat travel, light etc. Electromagnetic radiation is 1 way that energy travels through space.
Examples:
- Sunlight
- Xrays
- Radio, TV, sound waves
- Radiant heat (like from a campfire or stove)

- All show wavelike behavior
- All travel (in a vacuum) at \(3.00 \times 10^8\) miles/second

The study of waves is big – but there are:

3 characteristics to describe waves:

1. Wavelength (\(\lambda\) lambda) = distance between peaks and troughs

2. Frequency (\(\nu\) nu) = how many waves pass by a given point per second

3. Speed (\(v\)) = tells how fast a wave moves through space

The electromagnetic spectrum - Shows the spectrum of electromagnetic waves

- Light is one form of electromagnetic radiation
- Light appears to travel in waves
- But also behaves like a stream of tiny particles (called photons)
- Science currently describes light as having properties of both a particle and a wave
**Around 1912 – a Danish physicist –** Neils Bohr further refined our beliefs around atomic structure

- It was already known that elements (atoms) in gas state give off colored light if they are subjected to high temperature or voltage
- Each element emits energy with specific frequencies or wavelengths –
- When passed thru a prism or diffraction grating each element has a unique set of line spectrum

*From this:*

**Bohr:** theorized that electrons exist in specific regions at certain distances from the nucleus and that electrons orbit around the nucleus of an atom

**Planck** – believed that energy were emitted in packets called quanta – not as stream of photons

Therefore Bohr theorized that electrons have several energy levels or orbits at different distances from the nucleus

- An electron must be in 1 level, not between – therefore quantized
- As e- absorb energy, they jump to a higher energy level (= excited state)
- As e- release energy, they fall to a lower energy level, a quanta of light is emitted at a specific wavelength
- Lowest energy level = ground state
- Chemical properties of an element and its position on the periodic table depend on electron behavior within the atoms
- Most of our current knowledge about atoms is based on spectroscopy.

Bohr’s theory suggests:

1. Electrons have energy levels
2. Spectral lines result from radiation of small increments of energy (quanta) when electrons shift from one energy level to another

Bohr’s ideas did not work for heavier atoms

**Then French physicist Louis deBroglie** came up with another wave theory and Schrodinger (Austrian physicist) theory

1. Use for heavier atoms
2. Believed that all objects exhibit wave properties
3. Showed that could calculate the ‘probability’ mathematically of finding an electron in a certain region of the atom
=called Quantum mechanics in Physics

So...electrons are not moving in orbits, but are found in orbitals – a region in space around the nucleus where there is a high probability (90%) of finding an electron.

**So....Bohr’s theory:**

- Energy of e- is quantized, e- is restricted to certain energy levels

And...**Wave Mechanics Model**

- Predicts separate (discrete) principal energy levels in the atom

*This is what led to electron configurations that we use today (discussed later in this unit)*

d. **Discovery of the neutron and other subatomic particles**

In 1932, James Chadwick discovered the neutron

- The neutron has no charge (is neutral) and has a relative mass of 1 AMU
- Almost all chemical properties of matter can be explained in terms of the subatomic particles – protons, neutrons and electrons
- (Many other subatomic particles have been discovered since – such as the mesos, positrons, neutrinos etc... not sure of their role yet)

**The Nuclear Atom – More History...**

Discovery of positively charged particles came soon after the discovery of radioactivity by Henri Becquerel in 1896

Radioactive elements spontaneously emit radioactive particles from their nucleus:

- Alpha particles
- Beta particles
- Gamma rays
- Rutherford’s did experiments on these radioactive particles and thru his experiments Rutherford discovered that the nucleus is positively charged
- Also found that most of the atom consists of empty space – when we speak of the mass of the atom – we are actually referring to the mass of the nucleus as the rest of the atom (the electron orbitals) – add virtually no mass to the atom
- Combined mass of protons and neutrons in the nucleus accounts for 99.9% of the mass of any atom
- Electrons are small and account for only 0.1%
a. General arrangement of subatomic particles:
   - Rutherford’s experiment showed that the atom contained a dense and charged nucleus
   - Chadwick showed neutrons had mass but no charge
   - Rutherford found that light mass electrons offset positive charge of protons

*Based on this evidence, model of the atom proposed that it contained a nucleus surrounded by electrons*
B. Subatomic Particles, atomic number, atomic mass, and Isotopes

a. Subatomic particles:

Atoms contain subatomic particles:

- Electrons – negative charge in the orbitals
- Protons – positive charge in the nucleus
- Neutrons – no charge

![Diagram of subatomic particles]

<table>
<thead>
<tr>
<th>Particle</th>
<th>Symbol</th>
<th>Relative electric charge</th>
<th>Relative atomic mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>electron</td>
<td>e-</td>
<td>-1</td>
<td>0 (0.0005AMU)</td>
</tr>
<tr>
<td>proton</td>
<td>p</td>
<td>+1</td>
<td>1 AMU</td>
</tr>
<tr>
<td>neutron</td>
<td>n</td>
<td>0</td>
<td>1 AMU</td>
</tr>
</tbody>
</table>

Neutral atom: # of protons = # of electrons

b. Atomic number of the elements

- Atomic number = number of protons in the nucleus of an element
- Determines the identity of the element
- Shown in the periodic table

c. Isotopes of the elements

- **History:** Experiments proved that the mass of the atom had to be more than just the sum of protons and electrons – led to the concept of a neutron – had no charge but did have the same approx. mass as a proton
  **all atomic nuclei have neutrons except for the simplest hydrogen atom**
• **Isotopes - definition:** Isotopes = Atoms that have the same atomic # (same # of protons) but have different atomic mass are isotopes of that element

Therefore all atoms of the same element will have the same # of protons but may have a different # of neutrons therefore their atomic mass will be different

Therefore isotopes:

• Same # of electrons
• Same # of protons (=atomic#)
• BUT different number of neutrons in nucleus

• **Isotope notation:**

mass # \( \rightarrow \) A
atomic # \( \rightarrow \) Z E element

**i.e.** \( ^{12}_6 \text{C} = \text{carbon 12} \) \( ^{13}_6 \text{C} = \text{carbon 13} \)

\( ^{14}_6 \text{C} = \text{carbon 14} \)

**Remember:** mass # = # of protons + # of neutrons

Atomic # = # of protons

Number of neutrons = mass # - atomic #

• Most elements occur in nature as a mixture of isotopes –
• However, not all isotopes are stable – some are radioactive and decompose to form new elements

**i.e.:** there are 7 isotopes of oxygen of which 3 are stable; Oxygen 16, Oxygen 17, and Oxygen 18

**d. Atomic Mass**

**What is AMU?**

Atomic mass units represent relative mass of atoms as compared to carbon 12

• Carbon 12 was chosen as the standard for atomic mass – assigned a value of 12 AMU (atomic mass units)
Therefore 1 AMU = 1/12 mass of a carbon-12 atom = 1.6606 \times 10^{-24} \text{g}

Therefore in the table of atomic masses – all elements have AMU values that are relative to the mass assigned to the reference isotope C12.

**AMU’s of elements:**

- Since most elements occur as a mixture of isotopes with different masses the AMU represents relative mass of all the naturally occurring isotopes of that element.
- It is not a whole # because there is an unequal occurrence of the various isotopes of that element.
- Average atomic mass is calculated by multiplying the atomic mass of each isotope by fraction or percentage of each isotope present and then adding the results.

**Definition of atomic mass:**

- Atomic mass = is a weighted average that reflects the relative abundance of each isotope of that element, compared to atomic mass of carbon 12.
C. Energy levels and Electron configurations

Energy Levels

a. Energy levels of atoms
- These energy levels are designated by an n – n = positive #
- Lowest energy level is n=1
- as n increases, the energy of the e- increases and e- found (on average) further from the nucleus

b. Sublevels of energy levels
- each energy level has sublevels – the # of sublevels each energy level has is the same as the energy level itself. i.e., n=1 has 1 sublevel (=2 possible e-), n=2 has 2 sublevels (=4 possible e-), n=3 has 3 sublevels (=6 possible e-) etc.
- the sublevels are designated by a letter – the letter corresponds to the theoretical shape of the orbital in 3D. The probability of finding the e- in that space is ~90%

S orbital has max 2 electrons

P orbital has max 6 electrons

D orbital has max 10 e-

F orbital has max 14 e-
c. **Orbital Shapes:** (describes where the electrons can be found 90% of the time) – you don’t need to know these shapes, but it may help you to imagine what is being talked about when we refer to s, p, d, or f orbitals.

- s orbitals (1 of them to fill with 2 e-)—look like a sphere
- p orbitals (3 of them to fill with 2 e- each)—look like 2 lobes along an axis
  3 axis in all — x, y, z — designated as px, py, pz
- d orbitals (5 of them to fill)—note diagram — 4 leaf clover sort of...
- f orbitals (7 to fill) —

d. **Specific energy levels:**

- n=1, has 1 energy sublevel = 1s
- n=2, 2 energy sublevels = 2s, 2p
- n=3, 3 energy sublevels = 3s, 3p, 3d
- n=4, 4 energy sublevels =4s, 4p, 4d, 4 f
- ...

e. **Number of e- in each level/ sublevel**

- Electrons also have a property called ‘spin’ —can only spin in 2 directions
- Designated by ↑ or ↓
- e- in same orbital must have opposite spin — therefore each orbital can hold max 2 electrons, each spinning in opposite directions
- this rule called **Pauli Exclusion Principle**

therefore:

- 1s has 2 e-, each moving in opposite directions
- 2s has 2 e-
- 2p has 2e- for each of the 3 orbitals therefore = 6e-
- 3s has 2e-
- 3p has 6e-
- 3d has 10e-

As energy levels increase, # of sublevels increases because we are further from the nucleus and therefore more space for sublevels.

In an atom in the ground state, electrons are at their lowest energy level

However, e- can move to a higher energy level
Electron configurations

Electron Configurations in the Periodic Table

### a. Ways of writing/representing atomic structure:

1. Atomic structure diagrams (not common) *don’t use this one in this course* Shows nuclear makeup and electron structure of each sublevel

2. Orbital diagram *we don’t use this one in this course*
   - Show energy levels and orbitals as boxes
   - Shows spin of e- as for 2 e- in each orbit

3. Electron configuration *what we do in this course*
   
   Guidelines:
   - Only 2 e- for each orbital
   - e- occupy lowest energy levels first
   - Each sublevel contains 1 e- before a second electron enters

Lists principle energy level of orbital and number of electrons in each as an exponent
i.e., Fluorine $9e^-$

![Diagram showing electron configuration]

i.e., carbon $6e^-$

\[1s^2 2s^2 2p^2\]

i.e., silicon $14e^-$

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

b. **Valence electrons** – (highest energy level s and p orbital of the element)
   - Valence electrons = the electrons in the outermost (highest) energy level
   - These electrons are involved in bonding atoms to form compounds
   - If the electron is unpaired, it is available for bonding
   - If all paired, atom is stable – harder to form bonds (such as Noble gases)

**Examples**

a) $B = 2s^2 2p^1$ 3 valence e-, 3 available to bond

b) $N = 2s^2 2p^3$ 5 valence e-, 3 available to bond

c) $Na = 3s^1$ 1 valence, 1 available to bond

d) $Cl = 3s^2 3p^5$ 7 valence, 1 available to bond

c. **Full electron configuration**
   - You should be able to write the electron configuration from any point on periodic table for that element
   - Electrons must fill lower principle energy levels and orbitals first
   - Hint: If you know # of e- (from atomic #) you can begin filling from lower levels up

**Example:** $F = 9p = 9e^-$

Starting at H at 1s, fill lower energy levels until you reach 9 (adding up the superscripts) – go down by following periods as shown in above diagram – keep filling lower orbitals with electrons until you reach the total

\[F = 1s^2 2s^2 2p^5\]
Examples:

Mg has 12P therefore 12 e- so = 1s² 2s² 2p⁶ 3s²
S = 1s² 2s² 2p⁶ 3s² 3p⁴

- Or just use the periodic table and work your way down.

d. Filling orbitals, Period 4 down

Note diagram above, as of period 4 you will start filling d orbitals when you reach the transition elements

1. Transition elements

Example:

Cu = 1s² 2s² 2p⁶ 3s² 3p⁶ 4s² 3d⁹

- Period 4→ You go from 4s back to 3d at the transition elements
- For elements in the first 3 periods, the period # corresponds to the energy level and, it + the # of elements before it in that row, represent the # of valence e- of that element.
- d orbital = 1 less than s and p orbitals in that row (period)

2. Inner transition elements

- The 2 bottom rows represent the ‘inner transition elements’ – or the lanthanide and actinide series. = f orbitals
- f orbitals = 2 less than the s and p orbitals.

I’ll show examples once we cover abbreviated e- config – too long to do a full one!!

NB: the d and f orbitals represent lower energy levels and it is always the s and p orbitals for that element that represents the valence (outer orbital) of that element

e. Abbreviated electron configurations

For the elements lower on the periodic table, writing electron configurations is tedious therefore use abbreviated electron config.

1. On the periodic table, go back to the noble gas that comes just before the element you want to do.
2. Write the noble gas in square brackets i.e, [Ar]
3. Then list the electron config from the noble gas to the element
i.e.,

\[ P = 1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^3 \quad \text{or} \quad [Ne] \ 3s^2 \ 3p^3 \]

\[ Sn = 1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^6 \ 4s^2 \ 3d^{10} \ 4p^6 \ 5s^2 \ 4d^{10} 5p^2 \quad \text{or} \quad [Kr] \ 5s^2 \ 4d^{10} 5p^2 \]

Recap:

1. \# of period corresponds to highest energy level occupied by e- in that period
2. s and p of the period the element is in, are the valence e-
3. the d (1 less than the period) and f (2 less than the period) are in a lower energy level
4. the elements of a Group (=family) have the same outermost e-config except they are in a different energy level
5. there are some exceptions – which we won’t learn in this course – you just want to know they exist!

f. Chemical behavior

- We observe that the same orbitals occur on each level
- Groups of elements show similar chemical properties because of similarity of e- configuration of outer electrons (valence shell = same # for all members of that Group)
- Therefore periodic trends...
D. The Modern Periodic table and periodic trends

How is electron structure – shown as electron configurations– related to the periodic table?

a. The periodic table

- Periods = principal energy levels
  - =rows
  - Labeled 1 to 7
- Groups (Families) = columns in periodic table
  - Labelled as 1-18
  - Elements in these groups behave in a similar manner chemically
- Representative elements - Metals
  - Elements in Group 1,2, some in 13-15
  - Can have 1 type of charge
  - Includes alkali metals (Group 1), and alkali earth metals (Group 2)
- Representative elements - Non-Metals and semi-conductors
  - Elements in Group 13-18
  - Can have 1 type of charge
  - Includes Halogens (Group 17) and Noble gases (Group 18)
  - Semiconductors are between metals and non-metals – they are non-metals with some similar properties to metals – these include elements in the ‘stairway’ between the metals and nonmetals
- Transition elements
  - Elements in Group 3-12
  - May have more than 1 type of charge
  - Includes inner transition metals - Lanthanides and Actinides
    - Lanthanides are elements with atomic numbers 58-71 which fill 4f sublevel
    - Actinides are elements with atomic numbers 90-103 which fill 5f sublevel
Periodic Trends in Atomic properties

- The design of the periodic table is based on chemical and physical properties of the elements.
- It can be used to predict properties and reactions of a wide variety of substances.

Trends:

a. Ionization energy is the amount of energy needed to remove an e⁻ - if you think about it, who are the ones that give up e⁻ easily (the metals) so these have low ionization energy – the non-metals tend to take e⁻, so they hold tightly to their e⁻, so they have high ionization energy.
- As distance from the nucleus increases, IE decreases because there is not as much attractive force btwn e⁻ and nucleus.
- Some metals give up e⁻ easier than others.
- Nonmetals – higher IE than metals
  - Gain e⁻ to form anions
  - Most active = upper right – BUT not Noble gases.

b. Atomic radius
   
   Down the table
   - Increases in size as you go down the table.
- increase due to an increase in # of energy levels electrons can occupy (remember each period represents another energy level) – so distance from nucleus to outside edge of atom increases

Across the table
- decreases in size across table – as protons increase (even though electrons also increase the same amt) – pull of positive charge in nucleus increases, exerts more attraction on e-, pulls them closer to nucleus

c. Electronegativity (how strong is the atoms ability to keep e- close)
- When 2 different kinds of atoms share paired electrons unequally, one atom is partly positively and other partly negatively charged because 2 atoms exert unequal attraction for the electron pair.
- The attractive force that an atom of an element has for shared electrons in a molecule or polyatomic ion = electronegativity
- Different for each element

Example: HCl - Paired e- are closer to chlorine because more electromagnetic, therefore, that end partly negative and H end partly positive.

Electronegativity trend:

- Fluorine highest at 4
- Non-metals high – gain e-
- Metals low therefore lose e-
- Cs and Fr lowest at 0.7

- Generally electronegativity increases left to right across, and decreases down a group for the representative elements
- Higher electronegativity = the stronger an atom attracts electrons
- Polarity of a bond is determined by difference in electronegativity values of atoms forming bond.
- If electronegativity same = nonpolar covalent bond – e- shared equally
- If electronegativity very different = polar bond

Dipole – molecule that is electrically asymmetrical = charge oppositely at 2 points – creates the potential for an intermolecular attraction between 2 molecules...more later on this under London forces and hydrogen bonds...
**First a note on valence electrons before we continue with bonding ...**

The term ‘valence’ can be very confusing...there is a difference between ‘valence of an atom’ and ‘valence electrons’. For Chem 047 we only deal with the predictable elements, such as the representative elements, and therefore, the following explains the difference for these groups.

**Valence electrons** refer to the outermost energy levels when filling electron configurations for an element – the s and p orbitals. Remember, when doing an electron configuration, you always fill the orbitals of the lowest energy levels first. That means that both the s orbitals of the valence or outer shell will be filled before the p orbitals (this is called the Aufbau principle), resulting in a distribution of electrons that do not appear spaced out around the atom.

The following examples show a Lewis dot structure showing the distribution of electrons according to the filling of orbitals in an electron configuration. In a Lewis dot structure the 4 sides of an elements symbol represent the 4 outer valence orbitals – one side would represent the s orbital and the other 3 sides would represent each of the 3 p orbitals.

Boron has an electron configuration of $2s^2 2p^1$ so both electrons in the s orbital should be filled first and then the one electron in the first p orbital.

Silicon has an electron configuration of $2s^2 2p^2$; both the electrons in the s orbital are filled first followed by one electron in each of the first 2 p orbitals.

**However,** if we look at the arrangement of electrons around these same atoms in regards to bonding – it will appear to be different...

**Valence of an atom (=combining capacity or bonding capacity of an atom)** refers to the electrons that are available for bonding – the electrons available to bond are any unpaired electrons in the valence shell. Another thing to remember is that like repels like – in other words, electrons repel electrons (negatives repel negatives) – so electrons want to be as far away from each other as possible. So Boron and Silicon would look more like this:

![Lewis dot structures for Boron and Silicon](image)

So when the valence electrons are spaced out, it now appears that there are 3 unpaired electrons around boron (therefore, 3 bonding sites), and 4 unpaired electrons around silicon (4 bonding sites).

Long story short...when filling a Lewis structure to represent bonding, put 1 electron per side before starting around again...this is explained more thoroughly below.
E. Ionic and Covalent bonds and Lewis Structures:

a. Lewis structures
- Metals lose e- to form cation to get stable valence e- structure
- Nonmetals gain e- to form anions to get stable valence e- structure

For most elements – this stable config is 8 e-, 2s and 6p

Reminders:
- Atoms will rearrange e- structure to decrease chem potential energy (become stable) by losing, gaining or sharing e- with other atoms.
- Valence e- in outermost energy level forms chemical bonds
- Shown by Lewis structure (format that shows valence e- of an atom)

Lewis structure of elements
- Uses symbol of element for nucleus and dots for e-
- Paired dots = paired e- (*remember spin – 1 in 1 direction – 1 in the other – this is the pair*)
- Unpaired dots = unpaired e-
- Determine dot structure by determining the # of electrons – put pair for s orbital first then fill p orbitals (other 3 sides) with 1 dot per side first then fill.

Examples: B = boron = 2s² 2p¹

\[
\begin{align*}
B & : \quad \cdot \\
\text{Cl} & : \cdot \\
\text{Si} & : \cdot
\end{align*}
\]

Chlorine = Group 7A = 7 e-
Silicon = Group 4A = 4e-

b. Ionic Bond and Electron transfer
- elements want to attain stable amount of e- structure = 8 e- in outer energy level
- atoms that give up e- to become stable and positively charged (*more protons than e-*), form cations
- atoms that gain e- to become negatively charged, form anions
• if atoms react with one another, atoms will transfer e- to another atom – one becomes positively charged, other negatively charged
• these ions will then become attracted to each other by their opposite electrostatic charges
• **this force is called an ionic bond

**Lewis Dot of Ionic bond**

**Example**

\[ \text{MgCl} \]

Mg = 1s^22s^22p^63s^2 so valence e- = 3s^2 so will want to lose 2e- to become 2^+

Cl = valence e- = 7 (3s^2 3p^5) so will want to gain 1 e- to become 1-

Therefore, Mg has 2 to give, Cl takes only 1; therefore need two Cl for every Mg

\[ \text{Mg} + \text{Cl} + \text{Cl} \rightarrow \text{Cl}^- \text{Mg}^{2+} \text{Cl}^- \]

**Charged ions attract to form ionic bond**

• Cation – smaller than parent atom because it loses outermost e-, and remaining e- are attracted by greater # of protons ⊂ draw closer to nucleus
• Anion – larger than parent atom, nuclear attraction on each e- lessens because there are more e- than protons compared to parent atom ⊂ expands
• Ionic bond – attraction between oppositely charged ions
  Formed when e- transfer from 1 atom to another
• Metals have little attraction for their valence e-, ⊂ form ionic bonds when combined with nonmetals
• NB: outer energy levels containing 8 e- were formed in all negative ions – this is the result of pairing of all the s and p e- in these outer energy levels

**Example:** the formation of sodium fluoride (NaF) from its elements:
Example: the formation of Aluminum oxide ($\text{Al}_2\text{O}_3$) from its elements:

- Al needs to lose 3 electrons: $[\text{Ne}]3s^23p^1$
- O needs to gain 2 electron: $[\text{He}]2s^22p^4$

We will need to transfer 6 electrons.  

2 Al and 3 O are needed! 

\[
\begin{array}{c}
\text{Al} \\
\text{O}
\end{array} \quad \text{\rightarrow} \quad \begin{array}{c}
[\text{Al}]^{3+} \\
[\text{O}]^{2-}
\end{array}
\quad \text{or} \quad \text{Al}_2\text{O}_3
\]

Predicting formulas of Ionic Compounds

- In stable chemical compounds, each atom attains a noble gas configuration
- Chemical ionic compounds are electrically neutral: metal loses e-, nonmetals gain e-
- Electron structure is similar therefore, the elements in a Group form compounds with same atomic e- ratios

If you know one, you know the others in those Groups:

i.e, NaCl → if you know one, therefore,

- KCl
- LiCl
- RbCl
- CsCl

This works for representative metal Groups A1, 2A and 3 A.

Example

Formula for calcium sulfide is CaS and lithium phosphide is Li$_3$P.

So we can predict formulas for:

a) Magnesium sulfide MgS
b) Potassium phosphide K$_3$P
c) Magnesium selenide MgSe
c. **Covalent Bond and Electron sharing**

- Do not transfer e-
- Share e- between them
- Covalent bond = pair of e- shared between 2 atoms

**this is a predominant bond and makes “true” molecules - Exist as an aggregate of molecules not ions**

**Example 1:** Hydrogen – each atom has 1s electron

- Contributes one e- of pair shared jointly between 2 =hydrogen nuclei
- Overlapping 1s orbitals

H : H

- Probability of highest density in area between 2 nuclei
- Therefore nuclei shielded from each other and draw close together

**Example 2:** Chlorine – overlapping p orbitals

- One 3p orbital electron overlaps the other atom
- Attains stable structure of 8 e- by sharing an e- pair

Other examples: F, Br, I, O, N

**Ionic vs Covalent bonds – the 2 extremes:**

<table>
<thead>
<tr>
<th>Ionic</th>
<th>Covalent</th>
</tr>
</thead>
<tbody>
<tr>
<td>Atoms very different</td>
<td>Identical atoms</td>
</tr>
<tr>
<td>E transfer = charges ions</td>
<td>Shared e equally</td>
</tr>
<tr>
<td>bond = electronic attraction between charged ions</td>
<td>bond = mutual attraction of 2 nuclei for shared e-</td>
</tr>
</tbody>
</table>

These compare the extremes but there is a range of bonds between these 2 extremes which is an intermediate type of bond – on the covalent end it is a polar covalent bond
- Atoms not different enough to transfer electrons, but not similar enough to share paired e- equally – **this creates a Dipole**

Electronegativity indicates what type of bond in the range will result between 2 atoms.

d. **Using electronegativity to decide bond type:**

If difference between electronegativity is:

- > than 1.7-1.9 bond is more ionic then covalent
- <1.5 = covalent
- >2 =ionic

*must be careful to distinguish difference between polar bonds and polar molecules: a covalent bond between different kinds of atoms is always polar, but a molecule containing different kinds of atoms may or may not be polar, depending on its shape or geometry – dipoles can act in opposite directions and cancel each other out: i.e. linear tetrahedral

**Electronegativity Difference:**

<table>
<thead>
<tr>
<th>Electronegativity Difference</th>
<th>Bond Type</th>
<th>Electron Bonding</th>
</tr>
</thead>
<tbody>
<tr>
<td>0</td>
<td>Covalent</td>
<td>Electrons shared equally</td>
</tr>
<tr>
<td>0.4</td>
<td>Covalent polar</td>
<td>Electrons shared unequally</td>
</tr>
<tr>
<td>1.8</td>
<td>Ionic</td>
<td>Electron transfer</td>
</tr>
<tr>
<td>3.3</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

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2) connect atoms with dots (as paired e-) or dashes (as bonds)
   - H forms 1 bond
   - not bonded to each other except in peroxides, have 2 single bonds or one double bond

3) distribute remaining e- as paired electrons around each atom to give 8 e- (noble gas configuration)

4) if not enough for 8 e-/atom, form double or triple bonds
   - Double bond = 4 e- per atom
   - Triple bond = 6 e- per atom

***some exceptions

This gives bonding but not the shape of the molecules (*dealt with in a later course*)

**Example** how many valence e- in each of these atoms:

<table>
<thead>
<tr>
<th>Cl</th>
<th>H</th>
<th>C</th>
<th>O</th>
<th>N</th>
<th>S</th>
<th>P</th>
<th>I</th>
</tr>
</thead>
<tbody>
<tr>
<td>7</td>
<td>1</td>
<td>4</td>
<td>6</td>
<td>5</td>
<td>6</td>
<td>5</td>
<td>7</td>
</tr>
</tbody>
</table>

**Example** Write Lewis structure for water H₂O

![Lewis structure of H₂O]

**Example** methane CH₄

C = 4 valence e-, H = 1 valence e- each

Total = 8, H can only bond once therefore C must be central

![Lewis structure of CH₄]

Make paired e- between C – H pair

Usually single atom is central atom

**Example** carbon tetrachloride CCl₄

![Lewis structure of CCl₄]
Example CO$_2$

\[
\begin{array}{c}
\vdots \\
\text{O} \\
\vdots \\
\text{C} \\
\vdots \\
\text{O} \\
\end{array}
\]

Complex Lewis Structures

- There are some molecules and polyatomic ions where there is no single Lewis structure. (*not covered in this course*)

Sometimes one of the elements in a compound must be nice and share 2 of its electrons with another element in the compound, even if it has nothing to share in return – this is what happens in a triple bond:

Example: carbon monoxide$^1$ – carbon has 4 e$^-$ and oxygen has a total of 6e$^-$ = 10 e$^-$ total

Step 1: Find valence e$^-$ in all atoms. Add them together.
- C=4
- O=6
- Total=10

Step 2: Find octet e$^-$ for each atom and add them together.
- C=8
- O=8
- Total=16

Step 3: Gives you bonding e$^-$. Subtract step 1 total from step 2.
- 16-10=6e$^-$

Step 4: Find number of bonds by diving the number in step 3 by 2 (because each bond is made of 2 e$^-$)
- 6e$^-$/2= 3 bond pairs

Step 5: Find the number of nonbonding (lone pairs) e$^-$. Subtract step 3 number from step 1.
- 10-6= 4e$^-$=2 lone pairs

Use information from step 4 and 5 to draw the lewis structure.
Put both atoms side by side since there are only two atoms.

\[
\begin{array}{c}
\vdots \\
\text{C} \\
\vdots \\
\text{O} \\
\end{array}
\]

$^1$Explanation from: http://www.biochemhelp.com/lewis-structure-of-co.html
F. London Forces and Hydrogen Bonds:

**Note: all ionic compounds have ionic forces, and covalent compounds have London forces, polar covalent have dipole and/or hydrogen bonds and all polar molecules have dipole forces.**

Both London forces and Hydrogen bonds are a type of force that holds 2 atoms or molecules together; so unlike ionic and covalent bonds which hold a molecule together, London forces and hydrogen bonds are BETWEEN molecules.

Intermolecular forces have an important role in giving a substance its properties.

- Any molecule that is not symmetrical will have a permanent dipole – meaning one end will be slightly positive and one end will be slightly negative due to the ‘hogging’ of electrons by the one with more protons (more positives) – this makes the hogging end more negative (because the 𝑒⁻ spend most of their time near that atom) and the atom that does not hog the 𝑒⁻ is more positive (𝑒⁻ spend less time near this atom).

a. London forces²

- London forces are weak INTERmolecular forces between atoms or molecules that lie in close proximity with each other. Nonpolar molecules can form temporary dipoles due to the ever changing positions of the shared electrons. This means that all nonpolar molecules have London forces which act like weak dipole-dipole forces.

b. Hydrogen bonds

- Hydrogen bonds are a type of dipole force found BETWEEN 2 molecules that contain hydrogen and either N, O, or F. Hydrogen bonds are relatively strong dipole forces, and therefore, can influence properties such as melting point, boiling point, and solubility – it takes more energy to break a Hydrogen bond then some other dipole forces and London forces.

² Diagram from slideshare
Examples of hydrogen bonds

Note: $\delta$ this symbol means ‘slightly’

so $\delta^-$ means slightly negative, and $\delta^+$ means ‘slightly positive’ – referring to the charge of the two dipoles.

Water:

Ammonia: $\text{NH}_3$