A. Early Models of the Atom

1. The earliest models of the atom came in the 5th century B.C. when Democritus expanded on the ideas of his teacher and stated that the differences between substances were the direct result of differences in the size and shape of tiny, uniform, uncuttable particles.

2. In the 4th century, B.C., Aristotle rejected this idea and proposed that earthly matter had no properties itself. Instead, Aristotle proposed that matter was composed of four major elements in various proportions. These elements were water, air, fire, and earth.

![Diagram of elements]

It is important to note that Greek theories were based on pure thought. They were essentially philosophical and moralistic in nature and did little or nothing to suggest a direction for experimental work.

During the middle ages, a great deal of practical chemical knowledge accumulated from the investigations of Arab and European alchemists. This large body of experimental work contained information on the separation of metals from ores and numerous types of distillations. However, as with the Greeks, the ideas of the alchemists were mainly philosophical in nature.
3. In 1808, John Dalton reintroduced the idea of atoms and supported his atomic theory on firm experimental foundation. Dalton’s atomic theory states that

i) Elements are made up of extremely small particles called atoms.

ii) The atoms making up a particular element are all identical and different types of atoms have different properties.

iii) Each chemical compound is unique and consists of a particular combination of specific types of atoms put together in a distinctive way.

iv) Chemical reactions involve the reshuffling of the atoms in a compound to make new compounds. The new compounds are made from the same atoms which were present in the original compound.


**THE LAW OF DEFINITE PROPORTIONS**

Every pure sample of a particular compound always contains the same proportion by mass of the elements in the compound. (Explained by Dalton’s 2\textsuperscript{nd} and 3\textsuperscript{rd} hypotheses.)

**THE LAW OF MULTIPLE PROPORTIONS**

When different masses of one element combine with a specific mass of a second element, the mass ratios of the first element are small whole number ratios. (Explained by Dalton’s 3\textsuperscript{rd} hypotheses.)
THE LAW OF THE CONSERVATION OF MASS

The mass of the reactants equals the mass of the products.
(Explained by Dalton’s 4th hypotheses.)

In addition to the atomic theory, Dalton made a huge contribution to chemistry by showing how to calculate the atomic masses of the atoms involved in a reaction and how to find the number of each type of atom in the molecules. This allowed accurate analyses of compounds and the prediction of the amounts of each reactant needed to make a given product.

Assign 1-6

4. In 1897, J.J. Thomson discovered that atoms contained negatively-charged particles which he called “corpuscles” and were later named “electrons”. Later, he showed that atoms also contained positively-charged particles. Thomson proposed an arrangement for the positively and negatively charge particles inside an atom which was nicknamed “the plum pudding model”.

THOMSON MODEL OF THE ATOM

Atom consists of a ball of positive charge with negative charges distributed throughout the ball.
5. In 1911, Sir Ernest Rutherford performed the gold foil experiment in which a thin piece of gold foil was bombarded with alpha particles. [http://micro.magnet.fsu.edu/electromag/java/rutherford/](http://micro.magnet.fsu.edu/electromag/java/rutherford/). This experiment showed that the atom was mostly empty space but contained a dense positively-charge nucleus. Rutherford postulated that the nucleus contained all of the protons and most of the mass of the atom.

6. The “planetary” model of electron behaviour proposed by Rutherford suggested that electrons orbited the nucleus in the same way that the planets orbited the Sun. The problem was that the movement of a negatively-charge particle around a positively-charged particle would cause the electrons to radiate energy. This would eventually result in the electron spiraling into the nucleus.

[http://www.youtube.com/watch?v=yQP4UJhNn0I&list=FL1EQiVwpOQw-0Y4i7sd-ZsQ&index=2&feature=plpp_video](http://www.youtube.com/watch?v=yQP4UJhNn0I&list=FL1EQiVwpOQw-0Y4i7sd-ZsQ&index=2&feature=plpp_video)
Bohr's test.doc
7. In 1913, Niels Bohr came up with an equation that accurately predicted the pattern of energies that can be produced by a hydrogen atom. In order to derive his equation, Bohr suggested that electrons could only exist in fixed energy or “quantized” energy orbits. Since electrons could only exist in these orbits, it was not possible for them to spiral into the nucleus.

**BOHR MODEL OF THE ATOM**

Bohr proposed that the electrons in an atom are restricted to having certain specific energies and are restricted to following specific paths called “orbits” at a fixed distance from the nucleus. Electrons were only allowed to emit or absorb energy when they moved from one orbit to another.

Bohr’s model was very successful for hydrogen but it ran into problems because it could not be made to work for any atom having more than one electron.

Assign 7-12

hydrogen-atom.jar
B. Atomic Numbers and Atomic Mass

1. Chemical elements differ from one another by the number of protons in their nucleus.
   - H has 1 proton in its nucleus
   - He has 2 protons in its nucleus
   - Cl has 17 protons in its nucleus
   
   Conversely, any atom having 1 proton must be hydrogen, H
   - any atom having 2 protons must be helium, He
   - any atom having 17 protons must be chlorine, Cl

2. The ATOMIC NUMBER of an atom = the number of protons in the nucleus
   
   The ATOMIC NUMBER of an atom = the charge on its nucleus

   ![Diagram](image)

   In a neutral atom, there is no overall charge which means the number of positive and negative charges are the same

   - NEUTRAL ATOM
   - number of electrons = number of protons
When the number of protons and electrons are not equal the particle will have an overall charge and is referred to as an **ION**.

Adding negative electrons produce a **negative ion** while taking electrons away results in a **positive ion**.

### FOR IONS

\[
\text{number of electrons} = \text{protons} - \text{charge}
\]

Assign 13-17abc

The **ATOMIC MASS** of an atom is equal to the total number of protons and neutrons. The number of neutrons in an atom can be determined by subtracting the atomic number from the atomic mass.

\[
\text{number of neutrons} = \text{atomic mass} - \text{atomic number (protons)}
\]

If the values for the atomic mass and atomic number must be shown with the atomic symbol, the following super/subscript symbol is used.

This symbol is also written as

\[^{23}\text{Na} \quad (\text{since all sodium atoms have an atomic number of 11}) \]

\[\text{Na} - 23 \quad (\text{avoids having to write atomic mass as superscript})\]
### Example VIII.1: Determining Number of Subatomic Particles

<table>
<thead>
<tr>
<th>Problem</th>
<th>How many protons, electrons, and neutrons do Fe, Al$^{3+}$, N$^{3-}$ and $^{235}$U$^{2+}$ contain?</th>
</tr>
</thead>
</table>
| Solution | protons = 26  
            since Fe is neutral, electrons = protons = 26  
            neutrons = 56 – 26 = 30  

| 26  
| Fe  
| 55.8  |
| protons = 13  
since Al$^{3+}$, subtract 3 electrons, electrons = 13 – 3 = 10  
neutrons = 27 – 13 = 14  

| 13  
| Al  
| 27.0  |
| protons = 7  
since N$^{3-}$, add 3 electrons, electrons = 7 + 3 = 10  
neutrons = 14 – 7 = 7  

| 7  
| N  
| 14.0  |
protons = 92  

since $^{235}\text{U}^{2+}$, subtract 2 electrons, electrons = $92 - 2 = 90$

neutrons = $235 - 92 = 143$

Assign 18-19

3. **ISOTOPES** are atoms that have the same atomic number but different atomic masses. Since isotopes have the same atomic number, they have the same number of protons; however, since the atomic masses are different, they have different number of neutrons.

\[ ^{12}_6\text{C} \text{ has 6 protons and 12-6 = 6 neutrons} \]

\[ ^{14}_6\text{C} \text{ has 6 protons and 14-6 = 8 neutrons} \]

Assign 22 (ace...)

Most elements exist as a mixture of several different isotopes. The molar mass of element is an average value for the mixture of isotopes.
EXAMPLE VIII.2 | CALCULATING AVERAGE MOLAR MASSES

**Problem**
Chlorine exists as a mixture of 75.77% Cl-35 and 24.23% Cl-37. If the precise molar mass of Cl-35 is 34.968 852 g/mol and Cl-37 is 36.965 903 g/mol, what is the average molar mass of the chlorine atoms?

**Solution**

\[
\text{mass of Cl-35} = (0.7577) \times (34.968\,852\,g/mol) = 26.4959\,g/mol
\]

\[
\text{mass of Cl-37} = (0.2423) \times (36.965\,903\,g/mol) = 8.9568\,g/mol
\]

\[
\text{total mass} = 26.4959\,g + 8.9568\,g = 35.45\,g/mol
\]

If the exact masses of the isotopes are not given in the question, the atomic masses can be used instead.

\[
\text{Cl}^{35} = 75.77\% \quad \text{and} \quad \text{Cl}^{37} = 24.23\%
\]

\[
\text{average mass} = (0.7577 \times 35) + (0.2423 \times 37) = 35.485\,g/mol
\]

Assign 23-25abcd
C. **The Electronic Structure of the Atom**

1. When atoms are irradiated with energy, some of the energy is absorbed and then re-emitted. If the light emitted is passed through a prism and then onto photographic film, a **“LINE SPECTRUM”** is observed.

   ![Line Spectrum Diagram]

In 1913, Niels Bohr proposed a model, which explained the appearance of a hydrogen atom’s line spectrum. He proposed that the electron in a hydrogen atom could only exist in specific energy states. These energy states are associated with specific circular orbits which the electron could occupy around the atom. Electrons could move from one orbit to another by absorbing or emitting specific amounts of energy called a **“QUANTUM”** corresponding to the energy difference between orbits.

According to Bohr, the pattern of lines in the spectrum reflects the energy level pattern. The observed spectrum represents energy level differences occurring when an electron in a higher energy level gives off energy and drops down to a lower level.

Demo discharge tubes, wintogreen mints, flame emission and phosphorescence.

2. The **THEORY OF QUANTUM MECHANICS** emerged from Bohr’s theories of electron orbits. Several significant changes were made to the Bohr’s basic ideas, the most notable being that electrons occupied particular regions of space called orbitals depending on their energies instead of orbiting the nucleus along specific well defined paths.
An **ORBITAL** is the actual region of space occupied by an electron in a particular energy level.
3. The **lowest set of energy levels** for hydrogen is arranged as
follows. Each dash represents the energy possessed by a particular
orbital in the atom. The letters s, p, d, and f refer to four different
“types” of orbitals.

**ENERGY LEVEL DIAGRAM FOR HYDROGEN**

A **SHELL** is the set of all orbitals having the same n-value.
For example, the 3\(^{rd}\) shell consists of the 3s, 3p, and 3d orbitals.

A **SUBSHELL** is a set of orbitals of the same type.
For example, the set of five 3d-orbitals in the 3\(^{rd}\) shell is a subshell.

As can be seen on the above energy level diagram, all the orbitals
for a hydrogen atom with a given value of n have the same energy
(this is not true for atoms with more than one electron).
The rules governing which types of orbitals can occur for a given energy level, and how many orbitals of a given type can exist, are:

i) For a given value of “n”, n different types of orbitals are possible.

- for $n = 1$; only the s-type is possible.
- for $n = 2$; the s- and p-types are possible.
- for $n = 3$; the s-, p-, and d-types are possible.
• for n = 4; the s-, p-, d-, and f-types are possible.

ii) An s-type subshell consists of **ONE** s-orbital.
A p-type subshell consists of **THREE** p-orbitals.
A d-type subshell consists of **FIVE** d-orbitals.
An f-type subshell consists of **SEVEN** f-orbitals.
4. The energy level diagram for hydrogen must be modified to describe any other atom. The modified diagram below can be used for all **POLYELECTRONIC ATOMS** (atoms having more than one electron).

**ENERGY LEVEL DIAGRAM FOR POLYELECTRONIC ATOMS**

```
1s
2s
3s
4s
3p
4p
5s
6p
7s
6d
5d
4d
3d
```

Energy Levels for a Polyelectronic Atom
5. An **ELECTRON CONFIGURATION** is a description of how the electrons of an atom are arranged into orbitals. In particular, which orbitals in an atom contain electrons and how many electrons are in each orbital.

The addition of electrons to the orbitals of an atom follows 2 simple rules.

i) As the atomic number increases, electrons are added to the available orbitals. Electrons **are added to the orbitals having the LOWEST energy first**. The order in which orbitals are filled is

\[
1s \ 2s \ 2p \ 3s \ 3p \ 4s \ 3d \ 4p \ 5s \ 4d \ 5p \ 6s \ 4f \ 5d \ 6p \ 7s \ 5f \ 6d \ 7p \ ...
\]

ii) A maximum of **2** electrons can be placed in each orbital. This means there can be a **MAXIMUM** of:

- **2** electrons in an s-type subshell
- **6** electrons in a p-type subshell
- **10** electrons in a d-type subshell
- **14** electrons in an f-type subshell.
**EXAMPLE VIII.3 WRITING ELECTRONIC CONFIGURATIONS**

**Problem**
Write the electronic configurations for He, Li, O, and Cl.

**Solution**

<table>
<thead>
<tr>
<th>Element</th>
<th>Configuration</th>
</tr>
</thead>
<tbody>
<tr>
<td>He</td>
<td>1s$^2$</td>
</tr>
<tr>
<td>Li</td>
<td>1s$^2$ 2s$^1$</td>
</tr>
<tr>
<td>O</td>
<td>1s$^2$ 2s$^2$ 2p$^4$ or</td>
</tr>
<tr>
<td>Cl</td>
<td>1s$^2$ 2s$^2$ 2p$^6$ 3s$^2$ 3p$^5$ or</td>
</tr>
</tbody>
</table>

6. The following diagram shows how the periodic table can be used to predict electronic configurations.

![Periodic Table Diagram](image-url)
Silicon, Si, has 14 electrons,

- First 2 electrons fill the 1s orbital and complete 1st shell.
- Next 8 electrons fill the 2s and 2p orbitals completing the 2nd shell.
- Remaining 4 electrons fill the 3s orbital and partially fill the 3p orbital.

\[
\text{Si} = 1s^2 \, 2s^2 \, 2p^6 \, 3s^2 \, 3p^2
\]

or

\[
1s^2 \, 2s^2 \, 2p^6 \, 3s^2 \, 3p_x^1 \, 3p_y^1
\]

Technetium, Tc, has 43 electrons,

- First 2 electrons fill 1s orbital.
- Next 8 electrons fill 2s and 2p
- Next 8 electrons fill 3s and 3p
- Next 18 electrons fill 4s, 3d, and 4p
- Remaining 7 electrons fill 5s and partially fill 4d

\[
\text{Tc} = 1s^2 \, 2s^2 \, 2p^6 \, 3s^2 \, 3p^6 \, 4s^2 \, 3d^{10} \, 4p^6 \, 5s^2 \, 4d^5
\]

Or

\[
1s^2 \, 2s^2 \, 2p^6 \, 3s^2 \, 3p^6 \, 4s^2 \, 3d^{10} \, 4p^6 \, 5s^2 \, 4d_1^{4} \, 4d_2^{4} \, 4d_3^{4} \, 4d_4^{4} \, 4d_5^{4}
\]

Assign 26 a-e
7. The set of electrons belonging to a given atom can be divided into two subsets: the **CORE** electrons and the **OUTER** electrons.

The **CORE** of an atom is the set of electrons with the configuration of the nearest noble gas (He, Ne, Ar, Kr, Rn, Xe) having an atomic number less than that of the atom being considered.

The **OUTER** electrons consist of all electrons outside the core. Since core electrons normally don’t take part in chemical reactions, they are not always explicitly included when writing the electronic configuration of an atom.

Core notation is a way of showing the electron configuration in terms of the core and the outer electrons. In this notation, a portion of the electronic configuration is replaced by the symbol of the noble gas it represents.

\[
\begin{align*}
S &= \boxed{1s^2 \ 2s^2 \ 2p^6} \ 3s^2 \ 3p^4 \quad \rightarrow \quad S = [\text{He}] \ 3s^2 \ 3p^4 \\
He & \\
\text{Rb} &= \boxed{1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^6 \ 4s^2 \ 3d^{10} \ 4p^6} \ 5s^1 \quad \rightarrow \quad \text{Rb} = [\text{Kr}] \ 5s^1 \\
\text{Kr}
\end{align*}
\]

**Assign 27 a-e**

8. There are two exceptions to the configurations of elements up to Kr. Instead of finding

\[
\begin{align*}
\text{Cr} &= [\text{Ar}] \ 4s^2 \ 3d^4 \quad \text{and} \quad \text{Cu} = [\text{Ar}] \ 4s^2 \ 3d^9
\end{align*}
\]

the actual configurations are

\[
\begin{align*}
\text{Cr} &= [\text{Ar}] \ 4s^1 \ 3d^5 \quad \text{and} \quad \text{Cu} = [\text{Ar}] \ 4s^1 \ 3d^{10}
\end{align*}
\]
Filled or exactly half-filled d-subshells are especially stable

9. There are two rules for writing the electronic configurations for ions.

i) Negative Ions

To write the electron configuration of a negative ion, add electrons to the last unfilled subshell, starting where the neutral atom left off.

\[ O^{2-} = O + 2e^- \]
\[ O^{2-} = [He] 2s^2 2p^6 \]

ii) Positive Ions

- Starting with the neutral configuration, remove electrons from the outermost shell (largest n-value) first
- If there are electrons in both the s- and p-orbitals of the outermost shell, the electrons in the p-orbital are removed first

Write core notation for the atom, remove electrons in the order:

- p-electrons before s-electrons before d-electrons

\[ Sn^{2+} \rightarrow \text{Start with } Sn = [Kr] 5s^2 4d^{10} 5p^2 \]
\[ \text{remove } 2 \text{ e}^- \]
\[ Sn^{2+} = [Kr] 5s^2 4d^{10} \]
\[ \text{Sn}^{4+} \rightarrow \text{Starting with } [\text{Kr}] \ 5s^2 \ 4d^{10} \ 5p^2 \]

\[ \text{Sn}^{4+} = [\text{Kr}] \ 4d^{10} \]

**Assign 28 odd**

10. Valence electrons are electrons which can take part in chemical reactions.

Valence electrons are all the electrons in the atom except those in the core (full p) or in filled d- or f-orbitals

\[ \text{Al} = [\text{Ne}] \ 3s^2 \ 3p^1 \text{ has 3 valence electrons} \]
\[ \text{Pb} = [\text{Xe}] \ 6s^2 \ 4f^{14} \ 5d^{10} \ 6p^2 \text{ has 4 valence electrons} \]
\[ \text{Xe} = [\text{Kr}] \ 5s^2 \ 4d^{10} \ 5p^6 \text{ has 0 valence electrons (noble gas configuration)} \]

**Assign 29 odd**
D. Organizing the Elements — The Periodic Table

1. Following Dalton’s Atomic Theory, new elements were discovered with increasing frequency. By 1817, chemists had discovered 52 elements and by 1863 that number had risen to 62. Numerous attempts were made to organize the elements by their mass or by their chemical and physical properties but they had limited success.

2. In 1869 Russian chemist Dimitri Mendeleev published a method of organizing the elements according to both their masses and their properties. Mendeleev showed that when the elements are listed according to masses, certain properties recur periodically. He broke the list into a series of rows such that elements in one row were directly over elements with similar properties in other rows. He called each horizontal row a PERIOD and each vertical column a GROUP.

In certain cases, Mendeleev interchanged elements when their properties dictated that an element should be placed in a particular group in spite of contrary indications by its mass. Mendeleev also left gaps in his table for elements which he believed were not yet discovered. He was so confident in his method of organization that he made predictions of the properties of these undiscovered elements and of their compounds. When these elements were eventually discovered, they matched Mendeleev’s predictions quite closely. At last with Mendeleev’s Periodic Table, chemists had a
way to organize and understand their data, and predict new properties.

Assign 30 lab

3. As more and better data became available, chemists made a significant change to Mendeleev’s method of organizing the elements. The modern periodic table is organized according to **ATOMIC NUMBERS** rather than **ATOMIC MASSES**. This solved problems where different isotopic abundances caused the masses to be “out of order” for the elements Ar and K, Co and Ni, and Te and I. The Periodic Law summarizes the organization of the periodic table.

THE PERIODIC LAW

The properties of the chemical elements recur periodically when the elements are arranged from lowest to highest atomic numbers.

4. In the modern periodic table, a **PERIOD** is the set of elements in a given **row** going across the table. A **GROUP** or **FAMILY** is the set of all the elements in a given **column** going down the table.

There are several special groups, rows, and “blocks” of elements.

- The **REPRESENTATIVE ELEMENTS** are the “main groups” of elements.

- The **TRANSITION METALS** are the central block of elements which separates the two blocks of the representative elements.

- The **ALKALI METALS** are the elements in the first column (except hydrogen). Demo and http://www.youtube.com/watch?v=m5SkgyApYrY
• The **ALKALINE EARTH METALS** are the elements in the second column.

• The **HALOGENS** are the elements in group 17 headed by fluorine.

• The **NOBLE GASES** are the elements in group 18 headed by helium.

• The **LANTHANIDES** and **ACTINIDES** are the two rows below the main part of the table starting with lanthanum and actinium respectively.

Assign 31-33
5. Elements can also be classified according to their metallic character.

The properties of metals

- reflect light when polished (are shiny and have metallic lustre).
- are opaque.
- are good conductors of heat and electricity.
- are generally, but not always, flexible when in thin sheets.
- are generally malleable (can be hammered or rolled into sheets) and ductile (can be stretched into wires).
- are usually solid at room temperature with the exception of mercury.

The properties of nonmetals

- are gases, liquids, or brittle solids at room temperature.
- are poor conductors of heat and electricity.
- if solids, are dull to lustrous in appearance and opaque to translucent.
6. There are some elements that share some properties with metals and nonmetals. The nonmetals can be divided into two subgroups, those with very low electrical conductivities and those with fair to moderate electrical conductivities.

A **SEMICONDUCTOR** is a nonmetal having an electrical conductivity which increases with temperature.

Semiconductors were formerly called **metalloids** or **semimetals** because they have properties which resemble metals more than nonmetals. The important difference is that the electrical conductivity of metals **DECREASES** with increasing temperatures whereas the electrical conductivity of semiconductors **INCREASES** with increasing temperature.
7. There are two important trends in the periodic table which exists among the elements:

i) The properties of the elements change from metallic to nonmetallic going from left to right across the table.

ii) Elements become more metallic (or better metals) going down a family in the periodic table.

Assign 35-39

Periodic table lab
E. Chemical Bonding

1. An **ELECTROSTATIC FORCE** is a force existing as a result of the attraction or repulsion between two charged particles. All chemical bonding is based on the following relationships of electrostatics:

   - Opposite charge attract each other.
   - Like charges repel each other.
   - The greater the distance between two charged particles, the smaller the attractive (or repulsive) force existing between them.
   - The greater the charge on two particles, the greater the force of attraction (or repulsion) between them.

2. Each “period” on the periodic table represents a different layer of electrons; that is, a different “electron shell”:

   - The 1\(^{\text{st}}\) shell has 2 electrons and therefore the 1\(^{\text{st}}\) period has 2 elements.
   - The 2\(^{\text{nd}}\) shell has 8 electrons and therefore the 2\(^{\text{nd}}\) period has 8 elements.
   - The 3\(^{\text{rd}}\) shell has 8 electrons and therefore the 3\(^{\text{rd}}\) period has 8 elements.

(Note: for the purposes of this section, the transition metals, lanthanides, and actinides are **IGNORED**. Only the REPRESENTATIVE ELEMENTS will be considered.)
3. Going from left to right across a given period, the atomic number (and the number of protons) increases, and the positive charge on the nucleus increases. This increase in atomic number also brings an increase in the number of electrons surrounding the nucleus. All the electrons in a given shell can be assumed to have the same average distance from the nucleus. As the number of protons in the nucleus of the atom increases, there is a greater force of attraction for the electrons in the shell and the distance between the electrons and the nucleus decreases.

Assign 40-41

<p>| | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
</tr>
</tbody>
</table>

The size of atom (atomic radius) decreases as the atomic number increases within a period.
4. The shells surrounding the nucleus can be described as **OPEN** or **CLOSED**. An open shell is a shell containing less than its maximum number of electrons. A closed shell is a shell containing its maximum number of electrons. A **closed shell has the same number of electrons as a noble gas** while an open shell does not.

e.g. The 3rd shell, Na to Ar, can hold a maximum of 8 electrons: 3s²3p⁶. The atoms Na to Cl have less than 8 electrons in their 3rd shell so they are **OPEN**. The atom Ar has its outermost shell full with 8 electrons therefore it is **CLOSED**.

Assign 42-44 (a,c,e…)

Previously **VALENCE ELECTRONS** were described as all the electrons in an atom excluding those in the core or filled d- or f-subshells.

**VALENCE ELECTRONS** are the electrons in **OPEN SHELLS**.

Valence electrons are “reactable electrons”. The **NOBLE GASES** have NO valence electrons and are **NOT REACTIVE** but “F” and “Na” **HAVE** valence electrons and **ARE REACTIVE**.
5. Isolated atoms have their electrons placed in \( s, p, d, \) and \( f \) orbitals; however, when an atom is involved in a chemical bond some of the atom's orbitals are modified to allow electrons to be shared between adjacent atoms. Only VALENCE electrons are considered for bonding and the TRANSITION metals are ignored.

- There are a total of **FOUR** orbitals into which electrons can be placed (one “s” and three “p” orbitals).
- Each individual orbital holds up to **2** electrons.
- Since electrons repel each other, each electron added goes into a vacant orbital, if possible.
- Only after each orbital contains one electron will the addition of successive electrons require electrons to become “paired up”.

The following LEWIS STRUCTURES or “electron dot diagrams” show how the valence electrons are distributed in an atom.

![Lewis Structures]

Only “unpaired” valence electrons are available for bonding. “Paired” electrons usually do not react and do not take part in bonding.

The VALENCE (not valence electrons) of an atom = the number of unpaired electrons.
Valence is sometimes called **COMBINING CAPACITY**.

Assign 47

6. In order to form a positive ion, an electron must be removed from a neutral atom.

\[ \text{Li} + \text{energy} \rightarrow \text{Li}^+ + \text{e}^- \]

**IONIZATION ENERGY** is the energy required to remove an electron from a neutral atom. (The electron is removed from the outermost shell and is always a valence electron unless the atom has a closed shell.)

Ionization energy **INCREASES** left to right across a period since the nuclear charge increases. The noble gas at the end of any period will always have the highest ionization energy for that period because it has a “filled” or closed shell.

Ionization energy **DECREASES** top to bottom along a group since each shell is progressively further from the nucleus.

Ionization energy **INCREASES LEFT TO RIGHT** across a period and **DECREASES TOP TO BOTTOM** in a group.

Assign 48-51 (see lab for more depth)

Assign 53-56
F. Types of Chemical Bonding

1. Atoms can form ions by either gaining or losing electrons. Metal atoms generally form positive ions and nonmetal atoms form negative ions due to their difference in **electronegativity**. Electronegativity is the tendency of an atom to attract electrons from a neighbouring atom. Atoms of high electronegativity will take electrons (becoming negative ions) from atoms of lower electronegativity (becoming positive ions).

Assign 58-61

Electronegativity increases when the nuclear charge increases and decreases as the size of the atom increases.

| Electronegativity **INCREASES left to right** across a period and **DECREASES down** a group. |

In general, when an atom forms an ion, the atom loses or gains sufficient electrons to attain a closed shell.

\[
\text{Na (1 valence e\textsuperscript{-})} \rightarrow e^- + \text{Na}\textsuperscript{+} \quad (0 \text{ valence e}\textsuperscript{-}'s \text{ like Ne})
\]

\[
\text{O (6 valence e}\textsuperscript{-}'s) + 2e^- \rightarrow \text{O}\textsuperscript{2-} \quad (0 \text{ valence e}\textsuperscript{-}'s \text{ like Ne})
\]

The most common charges found when going across the periodic table are shown below. The elements in group 14 (C, Si, Ge, Sn, and Pb) are not included because C, Si, and Ge do not form simple ionic compounds and Sn and Pb are metals which most readily form +2 ions and only rarely form +4 ions.

<p>| Group | 1 | 2 | 13 | 14 | 15 | 16 | 17 | 18 |</p>
<table>
<thead>
<tr>
<th>Charge on ion</th>
<th>+1</th>
<th>+2</th>
<th>+3</th>
<th>-3</th>
<th>-2</th>
<th>-1</th>
<th>0</th>
</tr>
</thead>
</table>

2. An **IONIC BOND** is formed by the attraction of positive ions to negative ions. It is formed when an electron from one atom is **TRANSFERRED** to another atom thus creating a positive and negative ion.

\[
\text{Li}^+ \quad + \quad \text{F}^- \quad \rightarrow \quad \text{Li}^+ \quad \text{F}^-
\]

**IONIC BONDS** are formed when elements from opposite sides of the periodic table are combined; that is, when a **METAL** and a **NONMETAL** are combined.

Assign 57

**IONIC BONDS** are very **STRONG**, so that compounds held together by ionic bonds have **HIGH MELTING TEMPERATURES**.

In an **IONIC SOLID** there are no actual molecules. Instead, there is a matrix of alternating positive and negative ions in three dimensions. Ionic solids are described as **FORMULA UNITS** which are the lowest whole number ratio of positive to negative ions.
3. When an atom forms an ion, the resulting ion will be a different size than the corresponding neutral atom.

When an atom gains electrons to form a **NEGATIVE ION**, the nuclear charge does not change but since there are more electrons, the electrostatic repulsion increases and the ion becomes **LARGER** than the neutral atom.

![Cl to Cl-]

When an atom loses electrons to form a **POSITIVE ION**, the electrostatic repulsion between electrons decreases since there are fewer electrons. As such, the remaining electrons are pulled closer to the nucleus and the ion becomes **SMALLER** than the neutral atom.

![Na to Na+]

Assign 62-67

4. In a **COVALENT BOND** there is no electron transfer. Instead, the bond involves the equal sharing of electrons. A covalent bond is formed when two atoms having less than full shells of electrons are able to share one or more of their electrons with each other to attain filled electron shells.

The **OCTET RULE** states that atoms in groups 14 to 17 of the periodic table tend to form covalent bonds so as to have 8 **electrons** in their valence shells.

![Covalent bond between F and F to form F2]
Atoms that form covalent bonds have relatively high electronegativities. They attract each other’s electrons strongly but will not let go of their own electrons. This results in a “tug of war” and the electrons are shared in the bond.

**COVALENT BONDS** are formed when a **NONMETAL** combines with a **NONMETAL**.

Assign 68

5. Oxygen atoms are 2 electrons short of a full shell. By **sharing 4 electrons** and forming a **DOUBLE BOND**, the atoms can form a full octet.

![Oxygen Bond](image)

Similarly, nitrogen atoms are 3 electrons short of a full shell. By **sharing 6 electrons** and forming a **TRIPLE BOND**, that atoms can form a full octet.

![Nitrogen Bond](image)

Assign 69-71

6. In covalent bonds between two different types of atoms the electrons may not be shared equally between the two nuclei. This results in a **POLAR COVALENT** bond where one end of the bond is slightly more negative (δ-) and the other end slightly more positive (δ+). Chemical bonds can be classified according to their difference in electronegativities. The following table lists the electronegativities of the elements.
<table>
<thead>
<tr>
<th>Type of Chemical Bond</th>
<th>Electronegativity Difference</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ionic</td>
<td>1.7 or greater</td>
</tr>
<tr>
<td>Polar Covalent</td>
<td>between 0.2 and 1.7</td>
</tr>
<tr>
<td>Covalent</td>
<td>0.2 or less</td>
</tr>
</tbody>
</table>

\[
\text{NaCl} = |0.9 - 3.0| = 2.1 \text{ (ionic)}
\]

\[
\text{CH}_4 = |2.5 - 2.1| = 0.4 \text{ (polar covalent)}
\]

\[
\text{F}_2 = |4.0 - 4.0| = 0.0 \text{ (covalent)}
\]
7. The formula of covalently-bonded binary compounds can be predicted.

<table>
<thead>
<tr>
<th>Group</th>
<th>1</th>
<th>2</th>
<th>13</th>
<th>14</th>
<th>15</th>
<th>16</th>
<th>17</th>
<th>18</th>
</tr>
</thead>
<tbody>
<tr>
<td>Valence</td>
<td>1</td>
<td>2</td>
<td>3</td>
<td>4</td>
<td>3</td>
<td>2</td>
<td>1</td>
<td>0</td>
</tr>
</tbody>
</table>

If we want to predict that formula of a compound between N and F.

![Diagram](image)

Assign 72ace
G. Writing Lewis Structures

1. **LEWIS STRUCTURES** (electron dot structures) are used to help visualize the arrangement of bonds in molecules. The symbol is used to denote the nucleus and dots are used to indicate the number of valence electrons.

To write the Lewis Structure for an atom, write its chemical symbol surrounded by a number of dots which represent its valence electrons (electrons in outermost s and p orbitals).

\[
\begin{align*}
\text{Li} & \quad \text{Ca} & \quad \text{B} \\
\cdot & \quad \cdot & \quad \cdot \\
\cdot & \quad \cdot & \quad \cdot \\
\cdot & \quad \cdot & \quad \cdot \\
\end{align*}
\]

2. The Lewis Structure of an **ionic compound** is written by:

- determining the charge for each ion
- arranging the nonmetal ions symmetrically around the metal ion.

*e.g.* Draw the Lewis Structure for \(\text{MgCl}_2\)

\[
\begin{align*}
\text{Cl}^- & \quad \text{Mg}^{2+} & \quad \text{Cl}^- \\
\cdot & \quad \cdot & \quad \cdot \\
\cdot & \quad \cdot & \quad \cdot \\
\cdot & \quad \cdot & \quad \cdot \\
\end{align*}
\]

Assign 85
3. Drawing Lewis Structures of covalent compounds that obey the octet rule follow a simple set of rules.

- Count up the number of valence electrons in the molecule or ion. Add one electron for each negative charge and subtract one electron for each positive charge.

- Determine which atoms are bonded together and put 2 electrons into each bond.

- Use the remaining valence electrons to complete the octets of atoms surrounding the central atom(s) with the exception of “H” that only needs 2 electrons. Then place any remaining electrons, in pairs, on the central atom. (These non-bonding pairs of electrons are called LONE PAIRS.)

- If a central atom has less than an octet of electrons, have a neighbouring atom share electrons with the “deficient” atom by putting an extra pair (or pairs) of electrons into the shared bond.

- Tidy up, replace each pair of electrons involved in a bond with a dash, “—”.
Draw Lewis Structures for the following:

\[ \text{NH}_4^+ \]

\[ \text{CHO}_2^- \]

\[ \text{HOPO} \]

**Resonance Structures** exist when one Lewis Structure does not accurately depict the molecule because there are alternate diagrams for the same molecule as in \( \text{CHO}_2^- \) and HOPO.
4. There are a number of atoms that violate the octet rule. In addition to H, the atoms Be, B, and Al are exceptions to the tendency for covalently-bonded atoms to complete their octet. These atoms have such a low electronegativity that they can only gain one extra electron in a covalent bond for every electron they can contribute to the bond.

Be has 2 valence electrons and can only share 4 electrons (forming 2 bonds) while B and Al have 3 valence electrons and can only share 6 electrons (3 bonds).

The Lewis Structure for BF$_3$ is

```
\[ \begin{array}{c}
  :F & \equiv & B & \equiv & F \\
  :F & & & & \equiv \\
\end{array} \]
```

A molecule in which one or more atoms (other than hydrogen) does not possess a full octet of electrons is called an **ELECTRON-DEFICIENT** molecule.
5. Elements in the 3rd and 4th periods of the periodic table frequently attain more than an octet of valence electrons when they form covalent compounds. Other than the fact that the central atom will end up with more than 8 electrons, the same rules are used to draw Lewis Structures.

The Lewis Structure for PCl₅ is

![Lewis Structure for PCl₅]

Assign 86 a, c, e...
H. The Shape and Behaviour of Molecules

Lewis structures can be used to help visualize molecules in three dimensions. Since all electrons carry the same charge, it is reasonable to assume that electron pairs in bonds and lone pairs will be oriented in a molecule as far away from each other as possible. The valence electrons should be evenly spread out in regions of space around the central atom. This is the basis of the **VALENCE SHELL ELECTRON PAIR REPULSION theory** (VSEPR).

http://www.chem.arizona.edu/~jpollard/FlashComponents/VIP/vsepr/vsepr.html

molecule-shapes_en.jar

Summary of VSEPR shapes.

<table>
<thead>
<tr>
<th>Bonds</th>
<th>Lone Pairs</th>
<th>Shape</th>
<th>Example</th>
<th>Structure</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>0</td>
<td>Linear</td>
<td>BeCl₂</td>
<td><img src="image" alt="BeCl₂" /></td>
</tr>
<tr>
<td>3</td>
<td>0</td>
<td>Trigonal Planar</td>
<td>BCl₃</td>
<td><img src="image" alt="BCl₃" /></td>
</tr>
<tr>
<td>4</td>
<td>0</td>
<td>Tetrahedral</td>
<td>CH₄</td>
<td><img src="image" alt="CH₄" /></td>
</tr>
<tr>
<td>3</td>
<td>1</td>
<td>Trigonal Pyramidal</td>
<td>NH₃</td>
<td><img src="image" alt="NH₃" /></td>
</tr>
<tr>
<td>2</td>
<td>2</td>
<td>Bent(Non-Linear)</td>
<td>H₂O</td>
<td><img src="image" alt="H₂O" /></td>
</tr>
<tr>
<td>5</td>
<td>0</td>
<td>Trigonal Bipyramidal</td>
<td>PCl₅</td>
<td><img src="image" alt="PCl₅" /></td>
</tr>
<tr>
<td>3</td>
<td>2</td>
<td>T-Shaped</td>
<td>ClF₃</td>
<td><img src="image" alt="ClF₃" /></td>
</tr>
</tbody>
</table>
2. Polar bonds are a result of varying electronegativities among the elements. Since molecules usually possess more than one bond, the interaction of these bonds will determine the overall polarity of the molecule.

If the polar bonds of a molecule are oriented such that they are in opposite directions then the dipoles are canceled out and the molecule is nonpolar. If the polar bonds are oriented such that the compliment each other then the molecule will be polar.

As a general rule, a molecule will be polar if it has polar covalent bonds and there are lone pairs of electrons on the central atom. (Square planar molecules are an exception to this rule.)

Assign #10 page 201

3. Individual molecules are held together by covalent bonds between the atoms in the molecule. Such bonds are strong and are called intramolecular forces. There are also weak forces that hold individual molecules next to other molecules. These intermolecular forces are called van der Waals forces.

There are two main types of van der Waals forces:
• Dipole-dipole Force
• London Forces
Polar molecules are often referred to as **dipoles** because these molecules have a slightly positive and slightly negative end. As a result of these dipoles, polar molecules experience an attraction between the molecules, these forces are called **dipole-dipole forces** and they affect many of the properties of a compound such as boiling point.

![Dipole Diagram](image)

There is a special case of dipole-dipole forces known as hydrogen-bonding. A **hydrogen-bond** occurs in compounds that contain hydrogen bonded to highly electronegative nitrogen, oxygen, or fluorine. A hydrogen-bond is simply a particularly strong dipole-dipole force.

![Hydrogen Bond Diagram](image)

London Forces are the weakest of van der Waals forces and are the result of momentary dipoles.

![London Force Diagram](image)
London forces are the weakest type of bonding force known. In general, the more electrons an atom or molecule has, the stronger the London forces.

The greater the atomic number of an atom, the stronger the London forces it experiences.

**London forces are always present**, but are much weaker than covalent or ionic bonds. Hence, **London forces are important when they are the only force of attraction existing between two species.** That is, London forces are important between the following closed-shell species:

i) adjacent noble gas atoms, and

ii) adjacent covalently-bonded non polar molecules (made up of atoms having a full shell after bonding.)

Assign unit 8 73-78 (p180); unit 9 11-15 (p202)

Like dissolves like

\[
\text{NaCl} \rightarrow \text{Na}^+ + \text{Cl}^- \\
\text{CaCl}_2 \rightarrow \text{Ca}^{2+} + 2\text{Cl}^-
\]