Atoms and The Periodic Table

A. Early Models of the Atom

1. Democritus

   - Earliest model of the atom
   - Stated that the differences between substances were the direct result of differences in the size and shape of tiny, uniform, uncuttable particles.

2. Aristotle

   - Rejected this idea and proposed that earthly matter had no properties itself.
   - Proposed that matter was composed of four major elements, water, air, fire, and earth, in various proportions.

3. Arab and European alchemists

   - Concerned with wealth and power.
   - Transmutation, changing lead and other cheap metals into gold.
   - Invented many of the chemical methods of separation still used today.

4. In 1808, John Dalton

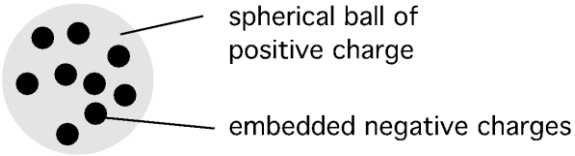
- Reintroduced the idea of atoms and supported his atomic theory on firm experimental foundation.
- Dalton's atomic theory states that
  - i) Elements are made up of extremely small particles called atoms.
  - ii) The atoms making up a particular element are all identical and different types of atoms 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.
- Model did not recognize electrical nature of the atom.

5. In 1897, J.J. Thomson

- Discovered that atoms contained negatively-charge particles which he called "corpuscles" and were later named "electrons".
- Later, he showed that atoms also contained positively-charged particles.
- Proposed "the plum pudding model".
- Problem with model because it required 100's of electrons per atom.

THOMSON MODEL OF THE ATOM

Atom consists of a ball of positive charge with negative charges randomly distributed throughout the ball.

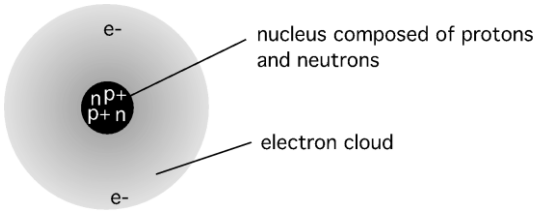


6. In 1911, Sir Ernest Rutherford

- Performed the gold foil experiment.
- Experiment showed that the atom was mostly empty space but contained a dense positively-charge nucleus.
- 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.

RUTHERFORD MODEL OF THE ATOM

Atom consists of a tiny, positively-charged nucleus surrounded by a cloud of negatively-charged electrons. The nucleus contains almost all of the mass of the atom and consists of protons and neutrons. The number of electrons surrounding the nucleus equals the number of protons in the nucleus, so as to make the atom electrically neutral.

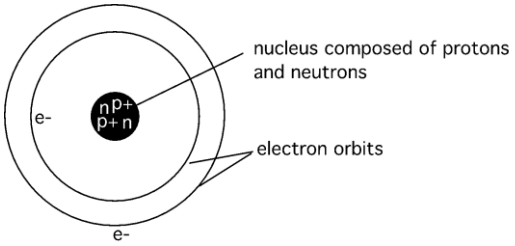


7. In 1913, Niels Bohr

- Came up with an equation that accurately predicted the pattern of energies that can be produced by a hydrogen atoms.
- 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.

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

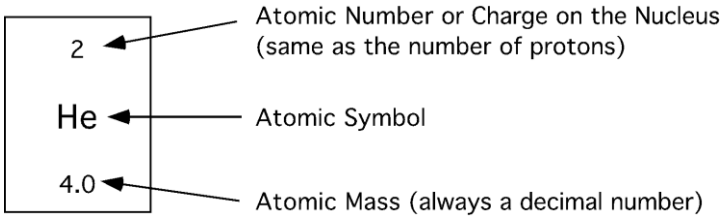
   Conversely, any atom having 1 proton must be hydrogen, H

   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



   In a neutral atom, there is no overall charge which means the number of protons and electrons are the same.

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

3. **Isotopes** are atoms that have the same atomic number but different atomic masses. Isotopes have the same number of protons but they have different number of neutrons.

$^{12}_{6}C$ has 6 protons and 12-6 = 6 neutrons

$^{14}_{6}C$ has 6 protons and 14-6 = 8 neutrons

The **mass number** of an isotope indicates the number of protons and neutrons.

Isotopes are written as:

^{23}Na (since all sodium atoms have an atomic number of 11) or as

Na – 23 (avoids having to write atomic mass as superscript)

Most elements exist as a mixture of several different isotopes. The **atomic mass** of element is a weighted average of the isotopes of an element.

4. Determining the number of subatomic particles for an atom:

   i. Protons = atomic number

   ii. Neutrons = mass number (or atomic mass) – protons

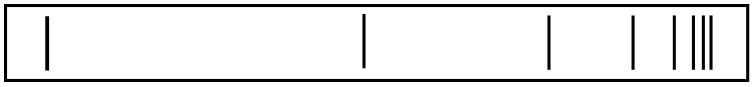
   iii. Electrons = protons – ionic charge

Symbol	Protons	Neutrons	Electrons
Fe			
Al^{3+}			
N^{3-}			
$^{236}U^{4+}$			
Kr			
As^{3-}			
Nb^{5+}			
$^{208}Pb^{4+}$			

EXAMPLE 8.1	**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	mass of Cl-35 = (0.7577) x (34.968 852 $^{g}/_{mol}$) = 26.4959 $^{g}/_{mol}$ mass of Cl-37 = (0.2423) x (36.965 903 $^{g}/_{mol}$) = 8.9568 $^{g}/_{mol}$ total mass = 26.4959 g + 8.9568 g = 35.453 $^{g}/_{mol}$ If the precise masses of the isotopes are not given in the question, the mass numbers can be used instead. ^{35}Cl = 75.77% and ^{37}Cl = 24.23% average mass = (0.7577 x 35) + (0.2423 x 37) = 35.485 $^{g}/_{mol}$

C. The Electronic Structure of the Atom

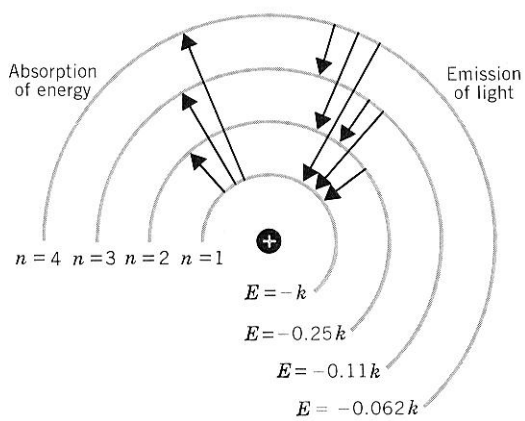
1. When atoms are irradiated with energy, some of the energy is absorbed and then re-emitted. If the light emitted is passed through a prism and then onto photographic film, a **"line spectrum"** is observed.



In 1913, Niels Bohr proposed that the electron in a hydrogen atom could only exist in specific energy states or orbits.

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.

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.



2. Quantum Mechanic Model (Wave Mechanics)

- Under appropriate circumstances, small bits of matter, such as electrons, behave like waves instead of particles.
- Electrons can have many different waveforms or wave patterns called **orbitals** which have characteristic energy.

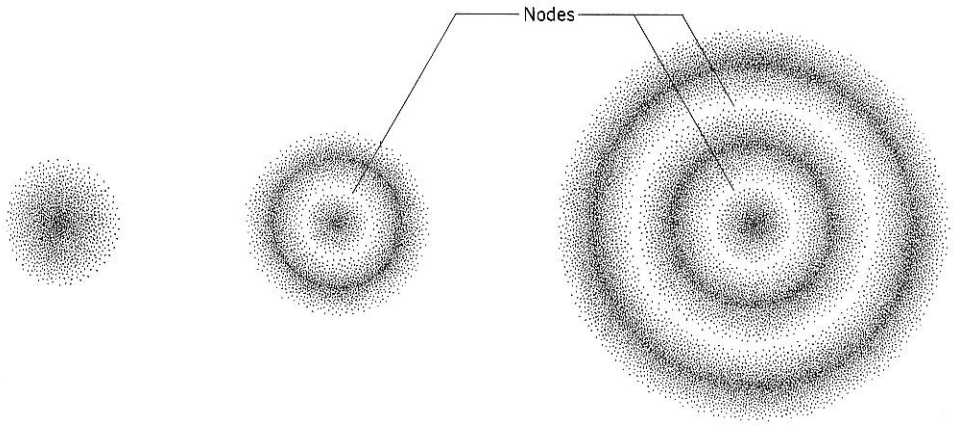
> An **orbital** is the actual region of space occupied by an electron in a particular energy level.

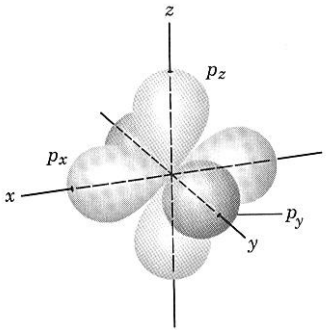
- Electron waves (orbitals) can be characterized by three quantum numbers, n, l, and m.
- n is called the principle quantum number, and all orbitals that have the same value of n are said to be in the same **shell**. The principal quantum number determines the size of the electron wave and hence how far the wave extends from the nucleus. n ranges from $n = 1$ to $n = \infty$.
- l is called the secondary quantum number, and it divides the shells into smaller groups of orbitals called **subshells**. The subshells are identified by the letters *s, p, d, f, g* ... l has values ranging from $l = 0$ to $l = n–1$.

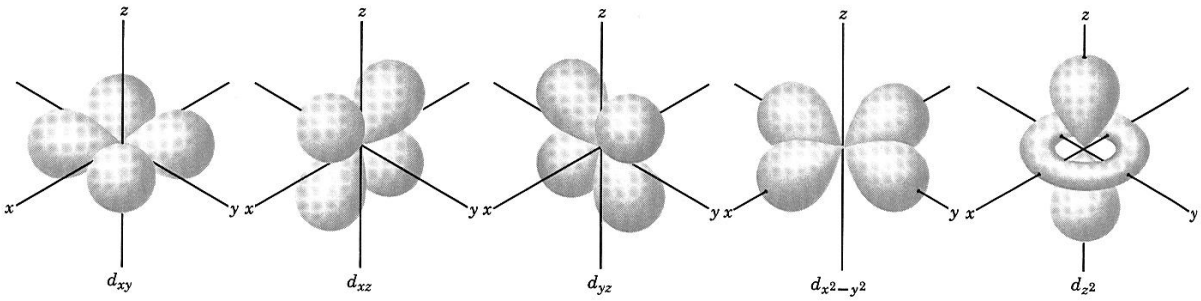
value of *l*	0	1	2	3	4
letter designation	*s*	*p*	*d*	*f*	*g*

- m is called the magnetic quantum number, and it splits the subshells into individual **orbitals**. This quantum number describes how an orbital is oriented in space relative to other orbitals. m has values ranging from $–l$ to $+l$.

s	*p*	*d*	*f*	*g*
1	*3*	*5*	*7*	*9*


Nodes

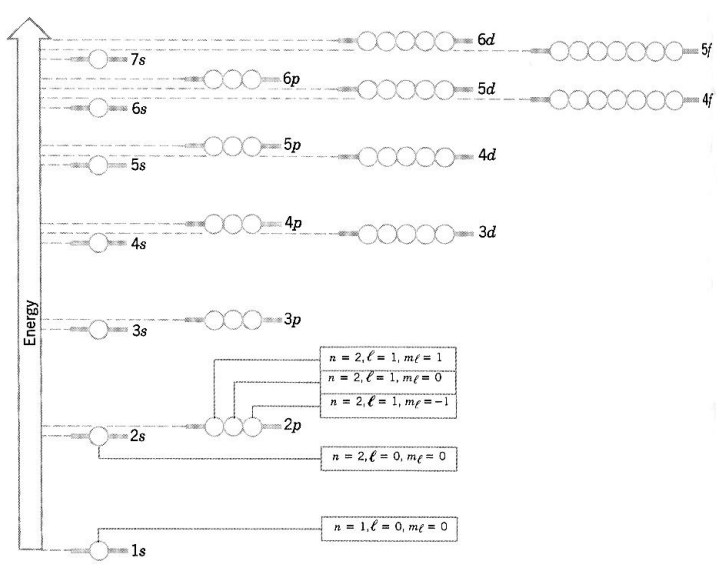

z
p_z
p_x
x
p_y
y


d_{xy}
d_{xz}
d_{yz}
$d_{x^2-y^2}$
d_{z^2}

Summary of Relationships among *n, l,* and *m*

Values of *n*	Values of *l*	Values of *m*	Subshell	Number of Orbitals
1	0	0	1*s*	1
2	0	0	2*s*	1
	1	–1, 0, +1	2*p*	3
3	0	0	3*s*	1
	1	–1, 0, +1	3*p*	3
	2	–2, –1, 0, +1, +2	3*d*	5
4	0	0	4*s*	1
	1	–1, 0, +1	4*p*	3
	2	–2, –1, 0, +1, +2	4*d*	5
	3	–3, –2, –1, 0, +1, +2, +3	4*f*	7

3. The relative energies of the orbitals is shown in the following diagram:



- Each orbital is represented by a separate circle – one for an *s* subshell, three for a *p* subshell and so forth.
- Also the all the orbitals of a given subshell have the same energy.
- In going upward on the energy scale, the spacing between successive shells decreases as the number of subshells increases. This leads to overlapping of the shells having different *n* values.

4. The **electron configuration** of an element describes how the electrons are arranged in an atom. In particular, the orbitals that the electrons occupy and the number of electrons in each orbital.

Electrons fill the orbitals starting with the lowest energy level first.

Subshell	Number of Orbitals	Maximum Number of Electrons
s	**1**	**2**
p	**3**	**6**
d	**5**	**10**
f	**7**	**14**

The following order of orbitals comes from the energy level diagram:

1s 2s 2p 3s 3p 4s 3d 4p 5s 4d 5p 6s 4f 5d 6p 7s 5f 6d 7p …

The order that orbitals are filled can be remembered using the following memory aid:

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

EXAMPLE 8.3	**WRITING ELECTRONIC CONFIGURATIONS**
Problem	Write the electronic configurations for He, Li, O, and Cl.
Solution	He has 2 electrons He = $1s^2$ Li has 3 electrons Li = $1s^2\ 2s^1$ O has 8 electrons O = $1s^2\ 2s^2\ 2p^4$ Cl has 17 electrons Cl = $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

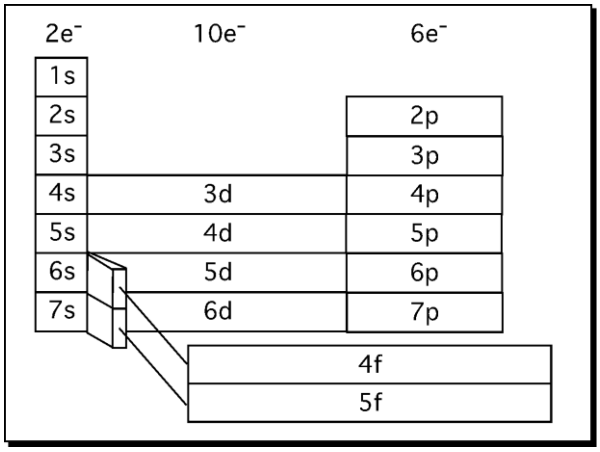
5. **Core notation** is a way of showing the electron configuration in terms of the core and the outer electrons. In this notation, a portion of the electronic configuration is replaced by the symbol of the noble gas it represents.

> The **CORE** of an atom is the set of electrons with the configuration of the nearest noble gas (He, Ne, Ar, Kr, Rn, At) 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.

$$S = \mathbf{1s^2 2s^2 2p^6} 3s^2 3p^4 \rightarrow S = [Ne]\ 3s^2 3p^4$$

$$Rb = \mathbf{1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6} 5s^1 \rightarrow Rb = [Kr]\ 5s^1$$

6. The periodic table can be used to predict the electron configuration of the elements:



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

$$Cr = [Ar]\ 4s^2\ 3d^4 \quad \text{and} \quad Cu = [Ar]\ 4s^2\ 3d^9$$

the actual configurations are

$$Cr = [Ar]\ 4s^1\ 3d^5 \quad \text{and} \quad Cu = [Ar]\ 4s^1\ 3d^{10}$$

Filled or exactly half-filled d-subshells are especially stable

8. 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.

$$O^{2-} = O + 2e^-$$

$$O^{2-} = [He]\ 2s^2\ 2p^4 + 2e^-$$

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

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

Sn^{2+} → Start with Sn = [Kr] $5s^2\ 4d^{10}\ 5p^2$

remove 2 e^-

Sn^{2+} = [Kr] $5s^2\ 4d^{10}$

Sn^{4+} → Starting with [Kr] $5s^2\ 4d^{10}\ 5p^2$

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

9. Atoms and ions with the same electronic configuration are said to be **isoelectronic**.

Ne = [He] $2s^2 2p^6$

O^{2-} = [He] $2s^2 2p^6$

Na^+ = [He] $2s^2 2p^6$

10. The electrons in the outermost shell of an atom determine the chemical reactivity of an element. Elements in the same group or family have similar electron configurations and hence have similar properties.

$$Li = [He]\ 2s^1$$

$$Na = [Ne]\ 3s^1$$

$$K = [Ar]\ 4s^1$$

The electrons in the outermost shell of an atom are called **valence electrons**.

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

$Al = [Ne]\ 3s^2\ 3p^1$ has 3 valence electrons

$Pb = [Xe]\ 6s^2\ 4f^{14}\ 5d^{10}\ 6p^2$ has 4 valence electrons

$Xe = [Kr]\ 5s^2\ 4d^{10}\ 5p^6$ has 0 valence electrons (noble gas configuration)

D. Organizing the Elements — The Periodic Table

1. In 1869 Russian chemist Dimitri Mendeleev published a method of organizing the elements according to both their masses and their properties.

   - When the elements are listed according to masses, certain properties recur **periodically**.
   - Elements were interchanged when their properties dictated that an element should be placed in a particular group in spite of contrary indications by its mass.
   - Gaps were left in his table for elements which he believed were not yet discovered. When these elements were eventually discovered, they matched Mendeleev's predictions quite closely.

2. The modern periodic table is organized according to **atomic number** rather than **atomic mass**. This solved problems where the masses appeared to be "out of order" for the elements Ar and K, Co and Ni, and Te and I.

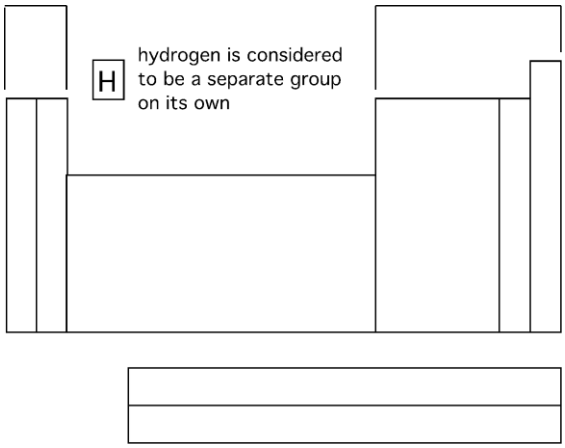
THE PERIODIC LAW

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

3. 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).
   - 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.



4. 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.

5. 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.

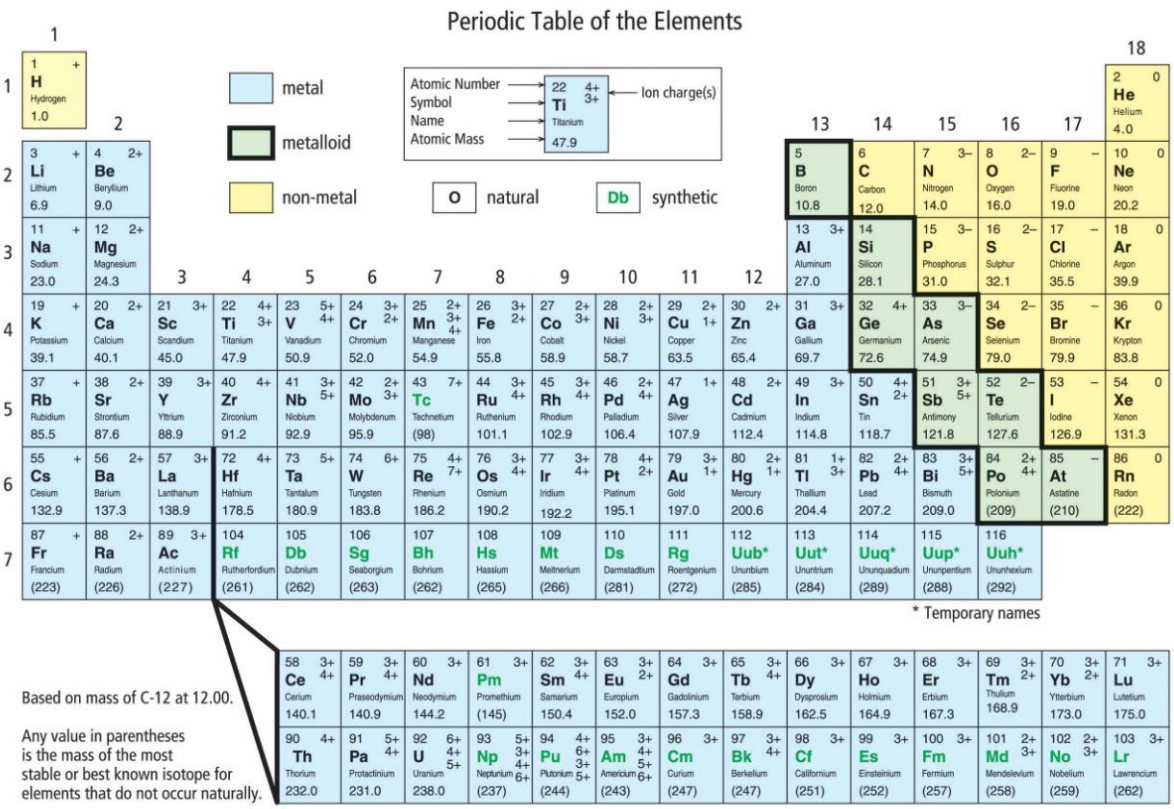


Figure 2.13 The periodic table of the elements

E. Periodic Trends

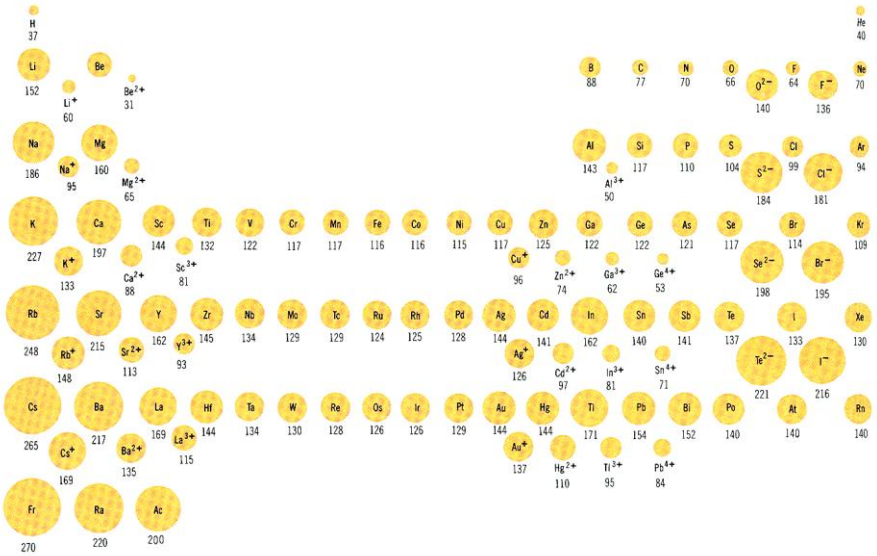
1. The periodic law states that when the elements are arranged according to increasing atomic number, certain properties recur periodically.

   This leads to the appearance of certain patterns or trends that can be predicted using the periodic table. Some of these trends include: metallic character, atomic radius (size), ionization energy and electronegativity.

2. **Metallic character** refers to how strongly an element behaves as a metal in terms of characteristic such as conductivity, malleability and ductility.

   - Elements become more metallic from right to left across a period.
   - Elements become more metallic from top to bottom within a group.

3. **Atomic radius** is the distance from the centre of the nucleus to the outermost shell.



- Atomic radius decreases from left to right across a period. 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.**

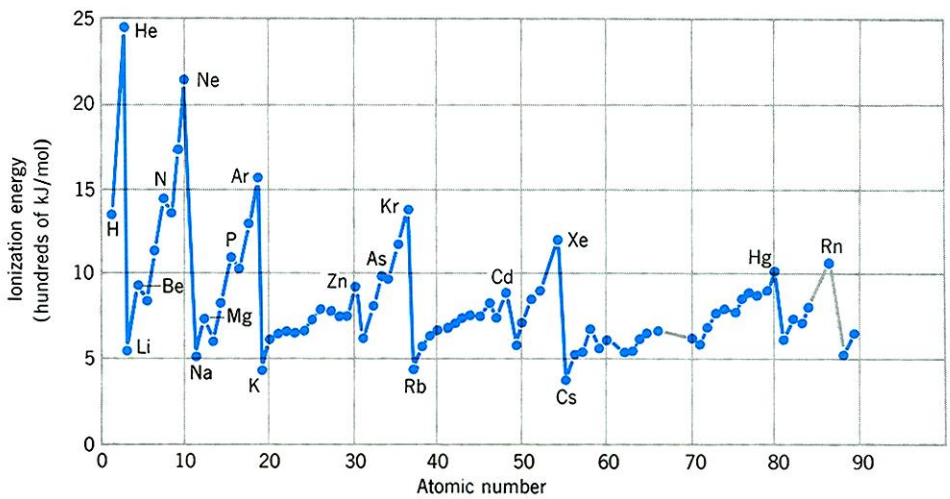
- As the number of protons in the nucleus of the atom increases, there is a greater force of attraction for the electrons in the shell and the electrons are pulled closer to the nucleus.

- Atomic radius increases from top to bottom within a group. Going down a group, the periods increase and hence the number of shells increase.

Atomic radius **decreases left to right** across a period and **increases top to bottom** down a group

4. **Ionization energy** is the energy required to remove an electron from the outermost shell of a neutral atom.

$$Li + energy \rightarrow Li^{+} + e-$$



- **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 which is very stable..
- **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.

5. **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).
- 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.

F. Types of Chemical Bonding

1. Atoms can form ions by either gaining or losing electrons. Metal atoms form positive ions and nonmetal atoms form negative ions due to their difference in **electronegativity**.

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

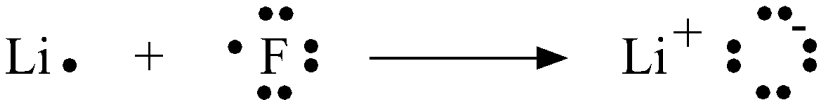
Na (1 valence e^-) $\rightarrow$ e^- + Na^+ (0 valence e^-'s like Ar)

O (6 valence e^-'s) + $2e^-$ $\rightarrow$ O^{2-} (0 valence e^-'s like Ar)

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.

Group	1	2	13	14	15	16	17	18
Charge on ion	+1	+2	+3	+4	-3	-2	-1	0

2. An **ionic bond** is formed by the mutual 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.

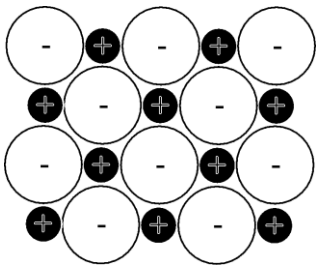


ionic bonds are formed when elements from opposite sides of the periodic table are combined; that is, when a **metal** and a **nonmetal** are combined.

ionic bonds are very **strong**, so that compounds held together by ionic bonds have **high melting temperatures.**

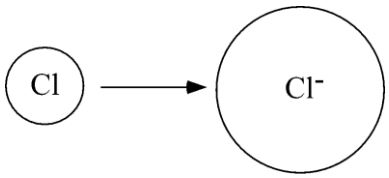
in an **ionic solid** there are no actual molecules. instead, there is a matrix of alternating positive and negative ions in three dimensions.

Ionic solids do not exist as molecules but are described as **formula units** which are the lowest whole number ratio of positive to negative ions.

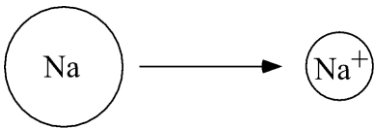


3. When an atom forms an ion, the resulting ion will be a different size than the corresponding neutral atom.

   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.

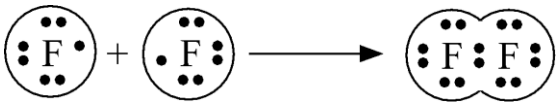


   When an atom loses electrons to form a **positive ion**, the electrostatic repulsion between electrons decreases since there are fewer electrons. As such, the remaining electrons are pulled closer to the nucleus and the ion becomes **smaller** than the neutral atom.



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.

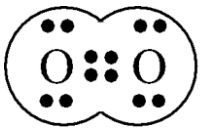
The **octet rule** states that atoms in groups 14 to 17 of the periodic table tend to form covalent bonds so as to have **8 electrons** in their valence shells.



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**.

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.



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.

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.

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17
2.1 H																
1.0 Li	1.5 Be											2.0 B	2.5 C	3.0 N	3.5 O	4.0 F
0.9 Na	1.2 Mg											1.5 Al	1.8 Si	2.1 P	2.5 S	3.0 Cl
0.8 K	1.0 Ca	1.3 Sc	1.5 Ti	1.6 V	1.6 Cr	1.5 Mn	1.8 Fe	1.8 Co	1.8 Ni	1.9 Cu	1.6 Zn	1.6 Ga	1.8 Ge	2.0 As	2.4 Se	2.8 Br
0.8 Rb	1.0 Sr	1.2 Y	1.4 Zr	1.6 Nb	1.8 Mo	1.9 Tc	2.2 Ru	2.2 Rh	2.2 Pd	1.9 Ag	1.7 Cd	1.7 In	1.8 Sn	1.9 Sb	2.1 Te	2.5 I
0.7 Cs	0.9 Ba	1.1 - 1.2 La-Lu	1.3 Hf	1.5 Ta	1.7 W	1.9 Re	2.2 Os	2.2 Ir	2.2 Pt	2.4 Au	1.9 Hg	1.8 Tl	1.8 Pb	1.9 Bi	2.0 Po	2.2 At
0.7 Fr	0.9 Ra	1.1 - 1.7 Ac-No														

Type of Chemical Bond	Electronegativity Difference
Ionic	1.7 or greater
Polar Covalent	between 0.2 and 1.7
Covalent	0.2 or less

$$NaCl = | 0.9 – 3.0 | = 2.1 \text{ (ionic)}$$

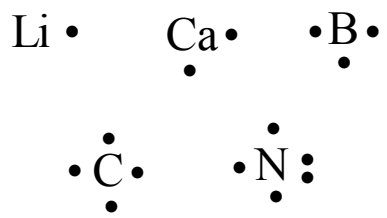
$$CH_4 = | 2.5 – 2.1 | = 0.4 \text{ (polar covalent)}$$

$$F_2 = | 4.0 – 4.0 | = 0.0 \text{ (covalent)}$$

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 indicate the number of valence electrons.

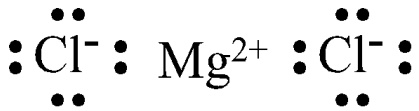
   To write the Lewis Structure for an atom, write its chemical symbol surrounded by a number of dots which represent its valence electrons (electrons in outermost *s* and *p* orbitals).



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 $MgCl_2$



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 negative 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.

Draw Lewis Structures for the following:

NH_4^+

H_2O

CO_3^{2-}

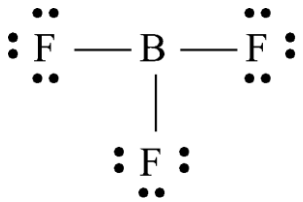
NH_3

Resonance structures exist when one Lewis Structure does not accurately depict the molecule because there are alternate diagrams for the same molecule as in CO_3^{2-}.

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).

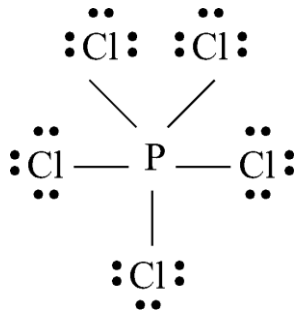
   The Lewis Structure for BF_3 is



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.

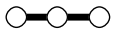
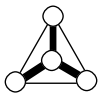
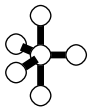
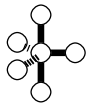
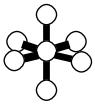
   The Lewis Structure for PCl_5 is

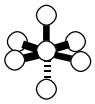
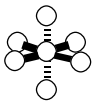


H. The Shape and Behaviour of Molecules

1. Lewis structures can be used to help visualize molecules in three dimension. 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).

   Summary of VSEPR shapes.

Bonds	Lone Pairs	Shape	Example	Structure
2	0	Linear	$BeCl_2$	
3	0	Trigonal Planar	BCl_3	
4	0	Tetrahedral	CH_4	
3	1	Trigonal Pyramidal	NH_3	
2	2	Bent(Non-Linear)	H_2O	
5	0	Trigonal Bipyramidal	PCl_5	
3	2	T-Shaped	ClF_3	
6	0	Octahedral	SF_6	

5	1	Square Pyramidal	BrF_5	
4	2	Square Planar	XeF_4	

2. Polar bonds are a result of varying electronegativities among the elements. Since molecules usually possess more that 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 **lone pairs of electrons on the central atom.** (Square planar molecules are an exception to this rule.)

3. Individual molecules are held together by covalent bonds between the atoms in the molecule. Such bonds are strong and are called **intramolecular forces.**

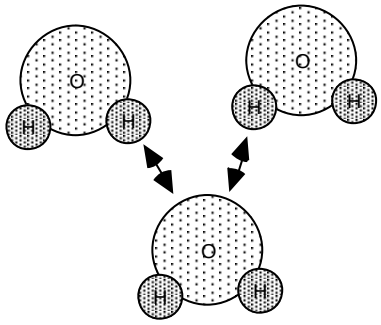
   There are also weak forces that hold individual molecules next to other molecules. These **intermolecular forces** are called **van der Waals forces.**

   There are two main types of van der Waals forces:

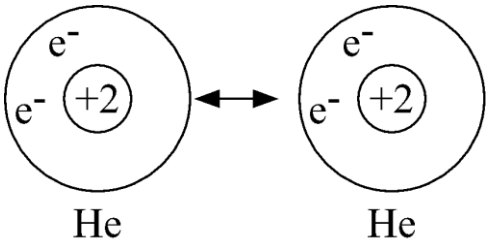
   - Dipole-dipole Force
   - London Forces

4. Polar molecules are often referred to as dipoles because these molecules have a slightly positive and slightly negative end. As a result of the 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.

5. 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.



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



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:

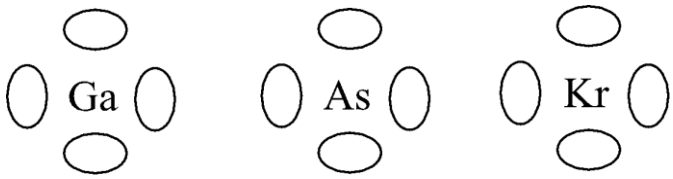
i) adjacent noble gas atoms, and

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

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.



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.**