The Periodic Table, **Electron Configuration** & Chemical Bonding

Lecture 7

Electrons

We will start to look at the periodic table by focusing on the information it gives about each element's electrons.

How electrons are configured about the nucleus is reflected in the structure of the periodic table.

The electron clouds of atoms are what interfaces when atoms meet, and dictate how they interact.

The periodic table

A Simple version with just a symbol and the atomic number of all the elements that make up matter.

The rows are called **periods** and the columns are called **groups**.

H 1																	He 2
Li 3	Be 4											B 5	C 6	N 7	O 8	F 9	Ne 10
Na 11	Mg 12											Al 13	Si 14	P 15	S 16	Cl 17	Ar 18
K 19	Ca 20	Sc 21	Ti 22	V 23	Cr 24	Mn 25	Fe 26	Co 27	Ni 28	Cu 29	Zn 30	Ga 31	Ge 32	As 33	Se 34	Br 35	Kr 36
Rb 37	Sr 38	Y 39	Zr 40	Nb 41	Mo 42	Tc 43	Ru 44	Rh 45	Pd 46	Ag 47	Cd 48	In 49	Sn 50	Sb 51	Te 52	I 53	Xe 54
Cs 55	Ba 56	La 57	Hf 72	Ta 73	W 74	Re 75	Os 76	Ir 77	Pt 78	Au 79	Hg 80	Ti 81	Pb 82	Bi 83	Po 84	At 85	Rn 86
Fr 87	Ra 88	Ac 89	Unq 104	Unp 105	Unh 106	Uns 107	Uno 108	Une 109	Uun 110	Uuu 111	Uub 112	Uut 113	Uuq 114	Uup 115	Uuh 116	Uus 117	Uuo 118

Lanthanide series	Ce 58	Pr 59	Nd 60	Pm 61	Sm 62	Eu 63	Gd 64	Tb 65	Dy 66	Ho 67	Er 68	Tm 69	Yb 70	Lu 71
Actinide series	Th 90	Pa 91	U 92	Np 93	Pu 94	Am 95	Cm 96	Bk 97	Cf 98	Es 99	Fm 100	Md 101	No 102	Lr 103

Periods and Groups

This table is set up in order of increasing atomic number.

It also has rows called **periods** and columns called **groups**. This structure gives the ground state **electron configuration** for any atom.

There are 4 quantum numbers associated with any electron that describe its existence in space, we will concern ourselves with two, one of energy level and that of its orbital.

principal energy level

There are 7 periods they are numbered 1-7 and represent the occupied **principal energy level**.

This is also called the principal quantum number and describes the distance of an electron from the nucleus.

As levels increase there are sublevels, or orbitals found within them. (s,p,d,f)

Rules

- Electrons should occupy the lowest energy states possible. (closest to the nucleus)
  - usually : s < p < d < f
- Energy states are denoted by principal energy level & sublevels determined by electron #.
- Sublevels and levels have a maximum number of electrons possible.
- When a sublevel fills the next level includes the previous.
- *Each element repeats the configuration of the previous and adds one electron.*

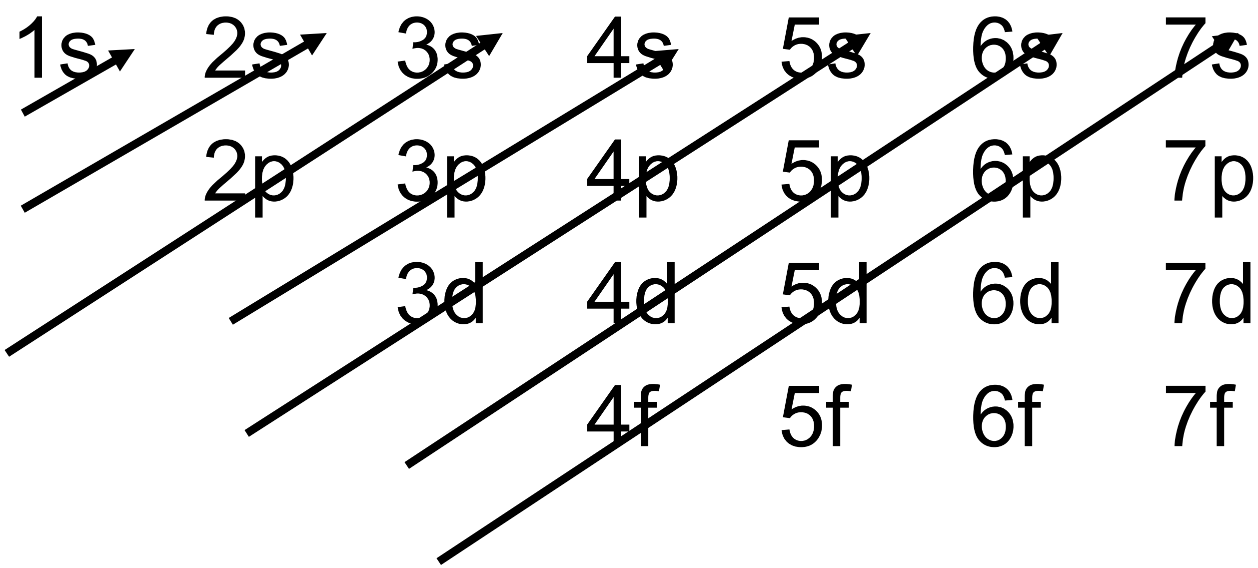
Higher levels

- We will leave the explanation of the filling of the other sublevels to more advanced courses.
- However, you may have noticed with potassium and calcium:
  - the d sublevel of principal energy level 3 does not start to fill until after the s sublevel of the next principal energy level fills (4s).

The Diagonal Rule for Configurations

Begin at the left and follow each arrow from tail to head and then from left to right.

There is an order to the filling of sublevels within a principal energy level. The filling of the levels is irregular.



Applying the Diagonal Rule

- Simply count electrons until the sublevel is filled, and then move to the next sublevel in the order given by the diagonal.
- Each new atom always repeats the pattern of the one before it.

Principal Energy Level	Sublevel(s)	Max. # of electrons	
1	*s*	2	level 1: up to 2
2	*s* *p*	2 6	level 2: up to 8
3	*s* *p* *d*	2 6 10	level 3: up to 18
4	*s* *p* *d* *f*	2 6 10 14	level 4: up to 32
n	*n* types of sublevels (*s*,*p*,*d*,*f* and others)	Level n up to $2n^2$ electrons	

What do these mean?

- $5f^8$
  - principal energy level? 5
  - sublevel ? f
  - Electrons in that sublevel? 8
- $3d^6$
  - principal energy level? 3
  - sublevel ? d
  - Electrons in that sublevel? 6

Notation

- Suppose the p sublevel on the principal energy level 3 contains 2 electrons.

We write this so: $3p^2$

How many electrons total does this atom have?

$1s^2 2s^2 p^6 3s^2 p^2$ = 14 electrons total

Element	Atomic number	Electron configuration
H	1	$1s^1$
He	2 level 1 complete	$1s^2$
Li	3	$1s^2 2s^1$
Be	4	$1s^2 2s^2$
B	5	$1s^2 2s^2 2p^1$
C	6	$1s^2 2s^2 2p^2$
N	7	$1s^2 2s^2 2p^3$
O	8	$1s^2 2s^2 2p^4$
F	9	$1s^2 2s^2 2p^5$
Ne	10 level 2 complete	$1s^2 2s^2 2p^6$

Noble Gases

- These items have 2 or 8 electrons in their outer principal energy levels.
- They are also found in the last column of the periodic table.
- The element following a noble gas is at the next higher energy level.
- To save space, the atom of that noble gas is written in brackets followed by the remaining configuration.

Orbitals

- These are parts of sublevels.

- They have 0, 1, or 2 electrons.

  0 is empty

  1 is half full.

  2 is filled.

sublevels

- s has 1 orbital.
- p has 3 orbitals. (it can hold up to 6 electrons)
- d has 5 orbitals.
- Orbitals are noted by dashes

___ ___ ___ ___ ___ ___ ___ ___ ___

s p d

Filling the orbitals

- Magnetic properties are visible and the electrons have different spin within an orbital.

___	empty orbital
↑	1 electron (+1/2 spin)
↑↓	2 electrons (+1/2 spin, -1/2 spin)

Orbitals fill in such a way that they go to the furthest point away within a sublevel, it half fills orbitals.

Element	sublevel	By orbital
H	$1s^1$	↑
He	$1s^2$	↑↓
Li	$1s^2 2s^1$	↑↓ ↑
Be	$1s^2 2s^2$	↑↓ ↑↓
B	$1s^2 2s^2 2p^1$	↑↓ ↑↓ ↑ ___ ___
		1s 2s 2p 2p 2p
C	$1s^2 2s^2 2p^2$	↑↓ ↑↓ ↑ ↑ ___
N	$1s^2 2s^2 2p^3$	↑↓ ↑↓ ↑ ↑ ↑
O	$1s^2 2s^2 2p^4$	↑↓ ↑↓ ↑↓ ↑ ↑
F	$1s^2 2s^2 2p^5$	↑↓ ↑↓ ↑↓ ↑↓ ↑
Ne	$1s^2 2s^2 2p^6$	↑↓ ↑↓ ↑↓ ↑↓ ↑↓

Valence electrons

- These are the outermost principal energy level electrons.
- The outermost principal energy level is the **valence level.**
- Valence electrons are only found in s & p sublevels.
  - They will vary between 1 and 8 electrons only.

Atomic Structure Determines:

- How molecules form & stay together
  - which atoms bond and in what quantity
- How molecules interact with other molecules and atoms

Types of Chemical Bonds

Covalent: electrons are shared between neighboring atoms, they orbit both nuclei holding the molecule together

Ionic: electrons move to electronegative atoms, creating charged atoms, oppositely charged atoms are held together by electrostatic force (+-)

(after the electron leaves an atom, the atom is positively charged, where the electron goes becomes negatively charged)

Metallic: such atoms have the ability to lose electrons creating a strong positive nucleus and the outer electrons move freely around all of the metallic atoms (making them good conductors)

Van der Waals: polar molecules interacting

(Polar molecules have partial charge)

- Like a Hydrogen bond between water molecules

For an example: Look at Water

- Chemical formula for water is H_2O
- Every water molecule has 2 hydrogen atoms and one oxygen atom
- How ?
- You must look at individual atomic structure to understand how the bonds formed

Bohr Model of the atom helps to visualize bonding

(A good starting point: not perfect though.)

- It is the electrons that interface between atoms, but proton charge is half of electrostatic force which holds molecules together
- Atomic # tells you proton # and electron #
- It is the amount of electrons in the outermost region that is of interest: **valence electrons** (usually equal to group or vertical column # of the periodic table)

The Atomic Number & Mass

- To know :
  - the # of **electrons** in an atom look at the atomic number, it is the same.
  - the # of **protons** in an atom look at the atomic number, it is the same.
    - (electron # and proton # are equal in the balanced stable state of an atom)
  - the # of **neutrons** in an atom subtract the atomic number from the mass
    - Total mass – proton mass = neutron mass

(Protons weigh just over 1 amu, neutrons just a bit more)

A **Neutron** weighs 1.009amu
A **Proton** weighs 1.007 amu

To calculate <u>neutrons</u> just round off and use 1.0 amu per proton & subtract it from the atomic mass.

Then round off the answer to equal the neutrons.

For example: Hydrogen (H)

- has an atomic number of 1
  - So it has 1 proton
  - And it has 1 electron
- has an atomic mass of 1.00794amu

  (neutrons & protons make up the mass)
- Its proton mass is (mass of protons) x (# of them)

  Atomic number

  1.007amu x 1proton = ~1amu=mass from protons

  Subtract it from the total mass of the atoms, to get the remaining mass from neutrons

  1.00794-1.007 = 0.00094 Neutrons = 0 N

How about an isotope of Hydrogen? the nuclide of deuterium : $^{2}_{1}H$

- has an atomic number of 1, just like all hydrogen
  - So it has 1 proton
  - And it has 1 electron
- **But it has a different atomic mass** of~2.0 amu

  (neutrons & protons make up the mass)
- Its proton mass is (mass of protons) x (# of them)

  Atomic number

  1.007amu x 1proton = ~1amu=mass from protons

  Subtract it from the total mass of the atoms, to get the remaining mass from neutrons

  ~2.0 amu - 1.007 = 0.993 Neutron = ~1 N

Try another: Sodium (Na)

- Atomic number is 11
- Atomic mass is 22.989768 amu

- Electron # ?
- Proton # ?
- Neutron #?

Sodium (Na)

- Atomic number is 11
- Atomic mass is 22.989768 amu

- Electron # 11
- Proton # 11
- Neutron # 22.989768 – 11 = 11.98
  - ~12 neutrons on average

Writing & Naming Chemicals

- Chemical formulas represent chemical entities (a single atom or molecules)
- The symbol and numbers are used to write them
- **Diatomic** molecules are made of 2 atoms:
  - H_2, O_2, N_2, F_2, Br_2, I_2, Cl_2 … (but not always the same type of element)
- **Polyatomic** molecules are made of 3 or more atoms: O_3,H_2O, CH_4…

Binary Compounds

- These are molecules with 2 elements
  - (but can have more than two atoms)

A few examples of ionic bonds:

You always start with the positive oxidation number bearing atoms

Na^{+} Cl^{-} = NaCl sodium chloride

Al^{3+} Cl^{-} $= AlCl_3$

You write this so $Al^{3+} + 3Cl^{-} = AlCl_3$