Scientists create models to understand how things work, including atoms.

Dalton created a theory for the atom with these 5 postulates

1. Elements consist of one or more tiny, indivisible particles called atoms
2. Atoms of the same element have the same properties
3. Atoms of different elements combine to form compounds
4. Compounds contain atoms in small, whole number ratios
5. Atoms can combine in more than one ratio to form different compounds

The later discovery of **radioactivity** (emissions from atoms) and **isotopes** (the same atom with different number of neutrons) showed postulates 1 and 2 of Dalton's theory to be false and led to the discovery of subatomic particles.

Subatomic Particles

	Location	Mass	Charge	Symbol
Proton	Nucleus	1	+1	p^+
Neutron	Nucleus	1	0	n^0
Electron	Outside nucleus	1/2000	-1	e^-

New theories of the atom, supported by experimental evidence, put protons and neutrons in a massive (meaning lots of mass), dense **nucleus** (or core) and placed the much less massive electrons in **shells** around the nucleus.

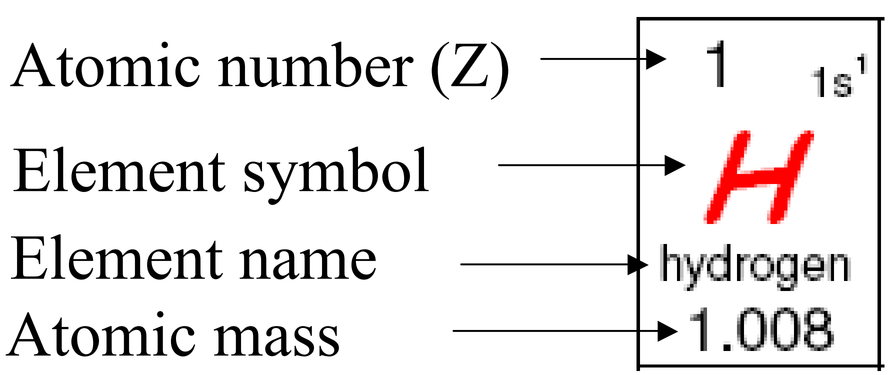
The periodic table lists elements by increasing atomic number, Z (see picture below from your periodic table handout – Week 1).

Atomic number is equal to

the number of protons in the nucleus of an atom

the number of electrons in a neutral atom

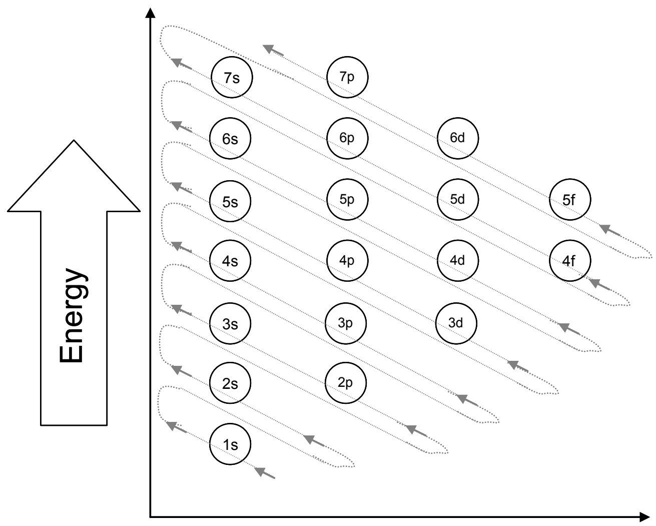
Hydrogen (Z=1) has 1 proton and 1 electron. Helium (Z=2) has 2 protons and 2 electrons.



The atomic mass of hydrogen (1.008) is the average value for all hydrogen's isotopes. If all hydrogen atoms had 1 proton and 1 electron, an atomic mass of about 1.002 would be expected. Because some hydrogen atoms have 1 neutron (deuterium) or 2 neutrons (tritium), the atomic mass is bigger than 1.002.

Quantum Theory

Quantum theory arranges electrons around the nucleus into energy shells and subshells. (This arrangement of electrons explains many properties of atoms, such as energy absorption and light emission that are beyond the scope of this course.) In the figure below (see handouts Week 1), the energy level is given by the first number in each circle, and the subshell is given by the second letter.



Electrons always fill starting in the 1s subshell. When that subshell is full, electrons start filling the next available subshell. Following the line back and forth through the subshells generates the complete filling order of subshells by electrons as shown below (this filling order is used below):

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

Quantum theory sets rules for how many electrons can fit into each subshell as follows:

s subshells hold a maximum of 2 electrons
p subshells hold a maximum of 6 electrons
d subshells hold a maximum of 10 electrons
f subshells hold a maximum of 14 electrons

To write the shorthand notation for the electron configuration (EC), indicate the number of electrons occupying each subshell with a superscript after the subshell. The EC shorthand $1s^1$ means one electron in the 1s subshell and a total of 1 electron in an atom (Hydrogen). If you look at the periodic table, the EC for hydrogen appears in the upper right hand corner of the element box.

Two electrons can fit in an s subshell. The EC shorthand $1s^2$ means two electrons in the 1s subshell and a total of 2 electron in the atom (helium). If you look at the periodic table, the EC for helium appears given in the upper right hand corner of the element box. After helium, the 1s subshell is full.

Electrons for hydrogen and helium go into the 1s subshell according to the rules above. The next subshells that fill are the 2s and 2p. Looking at the periodic table, lithium has 3 electrons (Z=3). To write the full EC for lithium, follow the rules above giving 2 electrons in the 1s subshell and 1 electron in the 2s subshell: $1s^2\ 2s^1$. Note that the element box for lithium only has the EC for the outermost subshell ($2s^1$). See the table below for the ECs for the remaining elements in the second row of the periodic table (note that the EC for the outermost electrons appears in upper right corner of the element box):

Element	Full EC	
Be	$1s^2\ 2s^2$	(the 2s subshell is now full)
B	$1s^2\ 2s^2\ 2p^1$	
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	$1s^2\ 2s^2\ 2p^6$	(p subshells hold 6 electrons)

After Ne, the second row (period) of the periodic table is full. The next subshells to fill are 3s and 3p (which can be seen using the filling order chart above or by looking at the element boxes for Na and Al). The full ECs appear below for elements in the third row:

Element	Full EC	
Na	$1s^2\ 2s^2\ 2p^6\ 3s^1$	
Mg	$1s^2\ 2s^2\ 2p^6\ 3s^2$	(the 3s subshell is full)
Al	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$	
Si	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$	
P	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$	
S	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$	
Cl	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$	
Ar	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$	(the 3p subshell is full)

After Ar, the third row (period) of the periodic table is full. (which can be seen using the filling order chart above or by looking at the element boxes for K and Sc, and Ga). The full ECs appear below for elements in the fourth row:

Element	Full EC	
K	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$	
Ca	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$	(the 4s subshell is full)
Sc	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^1$	
Ti	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^2$	
V	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^3$	
Cr	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^4$	
Mn	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^5$	
Fe	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$	
Co	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^7$	
Ni	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^8$	
Cu	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^9$	

Zn	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}$	(the 3d subshell is full)
Ga	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^1$	
Ge	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^2$	
As	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^3$	
Se	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^4$	
Br	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^5$	
Kr	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6$	(the 4p subshell is full)

The above process can be continued to generate the ECs for the remaining elements on the periodic table.

Example: Provide the full electron configuration for cobalt.

Solution: From the periodic table, cobalt has Z=27, so the neutral element will have 27 electrons to be configured using the above rules. Using the above filling order chart and rules (or the periodic table), fill electrons until 27 electrons have been filled. (Note that the periodic table lists cobalt as $3d^7$ so you know the stopping point.) From the chart, the filling order for the subshells is 1s 2s 3s 3p 4s 3d (and we can stop here). The full electron configuration for cobalt after placing the electrons is $1s^2\ 2s^2\ 3s^2\ 3p^6\ 4s^2\ 3d^7$.

The Periodic Table (and Quantum Theory)

Periodic Law: elements, when they are arranged according to increasing atomic number, have physical and chemical properties that recur in a regular pattern.

Groups (or families) are vertical columns of elements. The common names (when they exist) appear at the bottom of that column, such as the alkali metals for Group IA (except H), alkaline earth metals for Group IIA, halogens for Group VIIA, and noble gases for Group VIIIA.

1. Groups are divided into Group A (main group, representative, or s and p block elements) elements and Group B (transition or d block elements) elements.
2. The inner transition elements (or rare earth elements) include the lanthanides and actinides and do fit into the periodic table after the elements lanthanum (La) and actinium (Ac). These are the f block elements. To fit the periodic table on one page, they are pulled out and placed at the bottom.

Periods (or series) are horizontal rows of elements. There are 7 periods on the periodic table. Period1 has 2 elements (EC of 1s). Period 2 has 8 elements (ECs of 2s 2p), period 3 has 8 elements (ECs of 3s 3p). Periods 4 has 18 elements (ECs of 4s 3d 4p), and period 5 has 18 elements (ECs of 5s 4d 5p).

Trends in Physical Properties

Atomic Radius- decreases from bottom to top of the periodic table, and decreases from left to right across the periodic table. For example He is smaller than H, and Mg is smaller than Na.

Metallic Character - decreases from bottom to top of the periodic table, and decreases from left to right across the periodic table. For example, Ne is less metallic than Ar, and Mg is less metallic than Na.

Ionization Energy (the energy required to remove an electron from a gaseous element) - increases from bottom to top of the periodic table, and increases from left to right across the periodic table. For example Ne has a higher ionization energy than Ar, and Ca has a higher ionization energy than K.

Chemical Properties – elements in groups have similar chemical properties. For example, alkali metals form similar oxides (Li_2O, Na_2O, K_2O, and so on). Another example (working backwards) is if we know the chemical formula of calcium chloride is $CaCl_2$, then we can predict that the formula for magnesium chloride is $MgCl_2$. (Note that this relationship would be expected to hold between any alkaline earth metal and any halogen.)

1. The **noble gases**, or group VIIIA, are not chemically reactive. These elements have s and p subshells that are full. In other words, the valence shell (s and p subshells) are full.
2. **Metals** tend to lose electrons (and acquire a positive charge) to attain the electron configuration of the previous noble gas. For example, sodium (or Na) with an EC of $1s^2\ 2s^2\ 2p^6\ 3s^1$ would lose one valence electron, the $3s^1$, to form a positively charged ion and acquire the noble gas electron configuration of Ne, which is $1s^2\ 2s^2\ 2p^6$. To make writing ECs easier, ECs for elements can be written starting with the previous noble gas and then the EC for the remaining electrons. For example, the shorthand EC for sodium, Na is [Ne] $3s^1$. The chemical equation is $Na \rightarrow Na^+ + e^-$
3. **Nonmetals** tend to gain electrons (and acquire a negative charge) to attain the electron configuration of the next noble gas. Using the shorthand notation EC for phosphorous, P is [Ne] $3s^2\ 3p^3$, acquires 3 electrons to for an ion with three negative charges that looks like argon (Ar) and an EC of [Ne] $3s^2\ 3p^6$. The chemical equation is $P + 3e^- \rightarrow P^{3-}$
4. The **valence electrons** (or chemically reactive electrons) are always the electrons in the outermost s and p subshells. Sodium, Na, has one valence electron in the $3s^1$ subshell. Phosphorus, P, has 5 valence electrons located in the $3s^2$ and $3p^3$ subshells. The valence electrons are located farthest from the nucleus, and are therefore the most chemically reactive.
5. Note that the group number for Group A (the Roman numeral at the top of each column for Group A) gives the number of valence electrons directly. For example, Elements in Group IA have 1 valence electron. Elements in Group IIA have 2 valence electrons. Elements in Group IIIA have 3 valence electrons have 3 valence electrons, and so on. (Note that all of the transition metals, including the inner transition metals have 2 valence electrons using this classification because those electrons go into inner d and f subshells).

Example: Provide the electron configuration for the rhodium(II) or +2 rhodium ion.
Solution: Start by writing the full or shorthand electron configuration for rhodium (Rh, Z=45). Rh: [Kr] $5s^2$ $4d^7$. Forming the +2 ion means to remove two electrons from the outermost s and p subshells (valence electrons) giving an EC for the rhodium(II) ion as Rh^{+2}: [Kr] $4d^7$. Note that it was the 5s (valence) electrons that were removed first and not the 4d electrons.