

Chemistry 20

Lesson 2 – Atoms, ions and chemical bonds

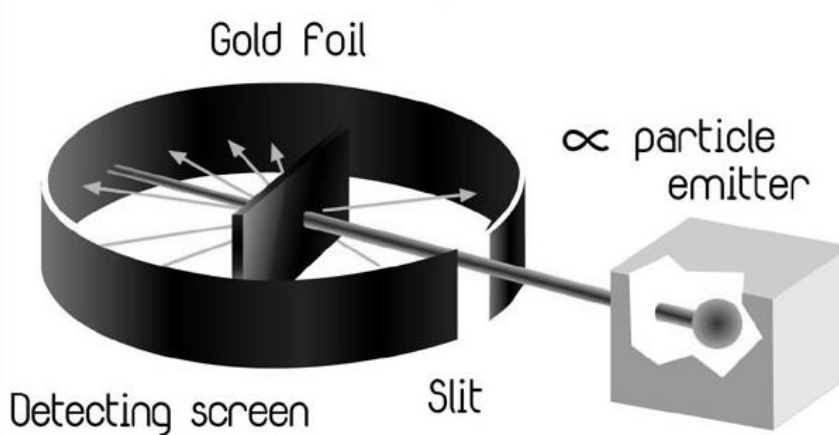
I. Atomic theories

As we saw in the previous lesson, Dalton believed that all elements were composed of extremely tiny, indivisible and indestructible atoms (i.e. solid spheres) and that all substances were composed of various combinations of these atoms. He also empirically verified that atoms of different elements were different in size and mass. For many years Dalton's idea worked quite well, however the results of further research began to contradict his ideas.

Following the work of a number of people (see Nelson pages 18 to 23) J.J. Thomson discovered the electron in 1902. His discovery indicated that atoms were themselves made of smaller particles called **subatomic particles**. In 1904, Thomson proposed a model of the atom that was based on the existence of subatomic particles. Thomson's model of the atom consisted of a sphere of positive charge and inside the positive sphere was an equal amount of negative charge distributed in the form of electrons. Thus the atom was neutral overall. He used the idea of raisins (electrons) embedded inside a large bun (positive sphere) as an analogy to the structure he imagined. Thomson argued that the chemical properties of the element might be associated with particular groupings of the electrons.

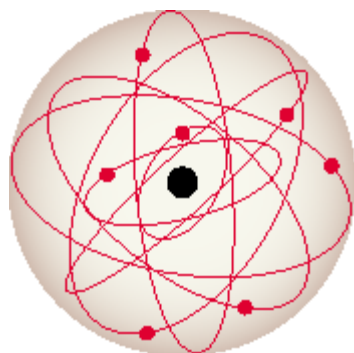
Ernest Rutherford (1871-1937) was a graduate student under J.J. Thomson and thus was strongly influenced by Thomson's work. In 1911, Rutherford began a series of experiments to verify the atomic model proposed by Thomson. The experiments were known as the **Gold Foil Scattering Experiments** and they had a tremendous influence on all atomic models from that point on.

In the experiment, gold foil was used as a target because it could be pounded down to a layer only a few atoms thick. Alpha (α) particles, which are actually helium nuclei, from a radioactive radium source were fired at the gold foil. Behind and around the gold foil was a zinc sulfide screen. Any particle hitting the zinc sulfide caused a glow of light which Rutherford could observe.



The results were **not** supportive of Thomson's model. While the vast majority of α particles went straight through the gold foil, as Thomson's model predicted, some were deflected at large angles. In fact, some particles were deflected straight back toward the source. Rutherford commented, "The result is as incredible as if you fire a 40 cm shell from a battleship at a strip of tissue paper and it reflects back at you." The Thomson model could not account for this result.

Rutherford proposed a model of the atom to try to account for the surprising results. He proposed a **nuclear** atom where the vast majority of the mass (99.98%) and all of the positive charge was concentrated in the central **nucleus** of the atom. Further, the nucleus accounts for only the smallest fraction, about 10^{-15} , of the atom's total volume – i.e. the atom was essentially empty space. Like a mini-solar system, the negative electrons orbited around the positive nucleus. This basic insight into the nature of an atom is still part of the current theory.



All matter is composed of atoms, and atoms are composed of three basic subatomic particles: protons, neutrons, and electrons.

Subatomic Particle	Relative Mass	Atomic Mass	Charge
proton	1836.12	1	+1
neutron	1838.65	1	0
electron	1	0	-1

The **nucleus** contains all of the **protons** and **neutrons** while the **electrons** are found around the nucleus.

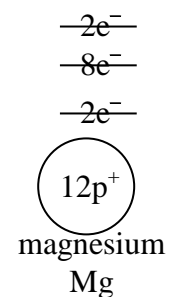
Although Rutherford's model was easy to visualize and understand, it had some serious flaws which we will not go into here. It was eventually replaced by **Bohr's model** of the atom and the **quantum mechanical** model which is our current model. In the quantum world, while we can never know how an electron is actually moving about the nucleus, we can say what its **energy** is and the **space where it will probably be found**. The electrons move as a "cloud" about the nucleus.

II. Arrangement of electrons – atoms and ions

The configuration or arrangement of electrons around the nucleus is based on the Bohr/Quantum model of the atom. In this theory electrons are located in certain specific energy levels about the nucleus called **shells**. Electrons within a shell have the same energy and each successive shell is further away from the nucleus. Scientists have discovered where the electrons are most often found in the three-dimensional space about the nucleus, but no one knows what type of path (if any) the electron is following. (The electron is not traveling around the nucleus like the planets travel around the sun.) However, experimental evidence and theoretical calculations have provided a strongly supported theory outlining a scheme of electron energy levels. There are seven possible energy shells and each shell can only hold a certain number of electrons.

Shell	1	2	3	4	5	6	7
Maximum # of e ⁻	2	8	8	18	18	32	32

To illustrate, let's consider a magnesium atom. In an **atom** the positive and negative charges are balanced – i.e. # of protons = # of electrons. An atom is **neutral**. Imagine a bare magnesium nucleus with no electrons initially. Magnesium nuclei contain 12 protons. Therefore a magnesium atom requires 12 electrons to form an atom. When we add electrons to the nucleus, the first 2 electrons form shell 1. Once a shell is full subsequent electrons occupy the next shell until it is full whereon electrons begin to populate the next shell and so on. The electron structure for a magnesium atom is shown to the right. The electrons in the outermost energy level are referred to as **valence** electrons.



Although the electron energy level theory was not derived from the development of the periodic table, there is a strikingly close relationship between the electron level representations for an atom and that atom's position in the periodic table. For instance:

1. The maximum number of electrons in each successive energy level equals the number of elements in each successive period (i.e. – there are 2, 8, 8, 18, 18, 32 and 32 elements in the seven periods of the periodic table. Compare this with the table above.)
2. The number of energy levels occupied by electrons equals the period number (i.e. – if an element is in Period 3, its atoms will have electrons in three energy levels).

III. Assignment (Part 1)

Using the magnesium atom electron energy diagram from the previous page as an example, complete the following abbreviated periodic table for **atoms**:

Group 1 Group IA								18 VIIIA
1		2 IIA	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	2
3	4	5	6	7	8	9	10	
11	12	13	14	15	16	17	18	

Questions:

1. What is the relationship between the old American system group number and the number of valence electrons?
2. What is the relationship between the period number and the number of energy levels in which electrons are accommodated?
3. What is the relationship between the maximum number of electrons in each energy level and the number of atoms in each period of the periodic table?
4. According to the above abbreviated periodic table, how many electrons can be accommodated before a new energy level is started in each of the first three energy levels?
1st energy level _____ 2nd energy level _____ 3rd energy level _____
5. Do the diagrams drawn above represent what the electron is actually doing? Explain.

IV. Arrangement of electrons – ions

In an **ion**, there are a different number of electrons than protons:

Atoms which **gain** electrons (# of e^- greater than # of p^+) become negatively charged ions called **anions**.

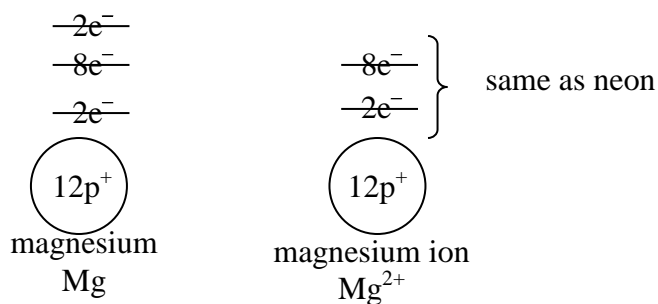
Atoms which **lose** electrons (# of e^- less than # of p^+) become positively charged ions called **cations**.

As we shall see below, in chemical reactions atoms may lose or gain electrons to acquire the more stable electron structure of the nearest noble gas. Atoms are **neutral** (zero net charge). Ions are **charged** (nonzero net charge). **Metals lose electrons to become cations** while **non-metals gain electrons to become anions**.

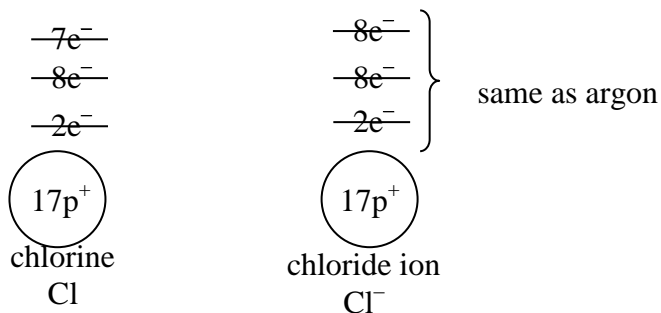
The energy level diagrams for simple ions in the first three periods are identical to the energy level diagrams for the closest noble gases. Atoms with 1, 2 and 3 valence electrons will lose their valence electrons to form cations with charges of $1+$, $2+$ and $3+$. For example, magnesium atoms will lose the two valence electrons when forming an ion. The net charge is:

$$(12+) + (10-) = 2+.$$

Thus we write a magnesium ion as Mg^{2+} .



Atoms with 5, 6 and 7 valence electrons will gain 3, 2 and 1 electrons to form ions with charges of $3-$, $2-$ and $1-$, respectively.



Atoms with 4 valence electrons in their second and third energy levels (i.e., carbon and silicon) do not form simple ions. These atoms tend to **share** their four valence electrons with other atoms.

Note:

1. All of the simple ions in the first three periods will have the same number of electrons as the nearest noble gas.
2. The name of a **nonmetallic ion** ends in *ide* while the name for a **metallic ion** uses the full name of the metal (i.e., chloride ion and magnesium ion).

- Hydrogen atoms may either gain or lose an electron to form 1+ or 1– ions.
- The quantum mechanical model of the atom explains the transition metal ions in terms of energy sublevels called orbitals.
- Correct “charge speak.” The charge on an ion like Ca^{2+} is always written 2+ (not +2) and should be verbalized as two positive or positive two but not plus two.

V. Assignment (Part 2)

Using the magnesium and chloride ion electron energy diagram from the previous page as examples, complete the following abbreviated periodic table for **simple ions**:

Group 1 Group IA							18 VIII A
1	2 IIA	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	
3	4			7	8	9	
11	12	13		15	16	17	

Questions:

- What relationship exists between the electron structure of a Group A ion and the electron structure of the nearest noble gas?
- Why do boron, carbon and silicon not form simple ions? How do they satisfy their electron requirements?
- What charge do the ions from the following groups assume?
Group 1(IA) Group (IIA) Group (IIIA) Group (VA) Group (VIA) Group (VIIA)
- What evidence is there that a noble-gas-like electron structure is stable?
- What are the differences in the chemical properties of a sodium atom and a sodium ion?

VI. Chemical bonds

If atoms could exist as singular entities they would. The noble gases for example, exist in nature as single atoms that are non-reactive. In other words, they do not require the presence of any other atoms to be stable and “happy”. They are so chemically inert that they were completely unknown to Mendeleyev when he made his table of elements in 1861. The first one, helium, was discovered in 1868.

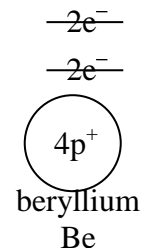
Other atoms, like fluorine, chlorine, cesium, and francium are extremely reactive. Why, then, are some elements reactive and others are not? First of all, a chemical interaction **involves the electrons only** – the nuclei of the different atoms are not affected. When atoms approach one another all of their electrons are simultaneously attracted by the nuclei of each atom. It is this attraction that causes electron rearrangements between two or more atoms to take place. Chemical reactions are the **rearrangement of atoms** into different organizations and forms. The reason they interact, according to quantum theory, is that all atoms attempt to meet two criteria:

1. To be **electrically neutral** (# of protons = # of electrons).
2. To acquire the more **stable electron structure of the nearest noble gas**.

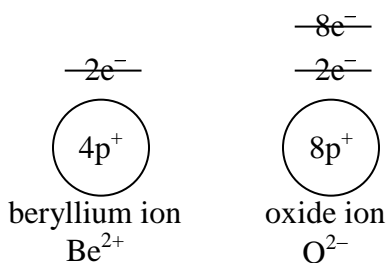
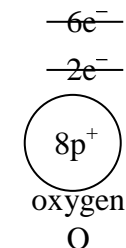
Chemical bonds are formed because most atoms can not satisfy these two criteria individually. To accomplish both criteria, atoms form relationships or bonds with one another. Two types of interactions or relationships are possible:

1. **An exchange of electrons** – where one atom gives up electrons to another atom. In other words, one atom loses electrons to become a cation while the other atom gains the electrons to become an anion. The loss and gain of electrons results in an **ionic bond** between the atoms.

For example, a beryllium atom has 4 electrons in 2 shells. While it is electrically neutral, it does not have an electron configuration like a noble gas. It can either gain 6 electrons to have a full 3rd shell (like neon) or it can lose 2 electrons and have a full 2nd shell (like helium). The easiest way is for beryllium to **lose** two electrons to become a Be²⁺ ion.



The second piece to the puzzle is to bring in another atom like oxygen which will accept or gain electrons. As we can see in the diagram to the right, an oxygen atom does not have an electron configuration like a noble gas. It can either lose 6 electrons to have a full 1st shell (like helium) or it can gain 2 electrons and have a full 2nd shell (like neon). The easiest way is for oxygen to **gain** two electrons to become an O²⁻ (oxide) ion.



Thus, if beryllium and oxygen form a relationship (a bond) they can solve each others needs. Beryllium needs to lose two electrons and oxygen wants to gain two electrons. Further, if they remain together they will be electrically neutral. Thus, beryllium and oxygen exchange electrons and form an **ionic compound** – beryllium oxide (BeO). Both atoms have the electron structure of the nearest noble gas and together they are neutral. Therefore, both criteria for “happy” atoms are met.

2. **Sharing of electrons** – where atoms share electrons with one another. This type of bond is referred to as **molecular (covalent) bonding**. An example of this kind of bonding is water, H_2O . Hydrogen has 1 electron and requires another electron to have a full outer shell. Oxygen requires two electrons to fill its outer shell. Since both atoms require more electrons, they form a bond by sharing electrons with one another. The electrons belonging to the hydrogen atoms spend part of their time around the oxygen atom, while two electrons from oxygen spend part of their time around the hydrogen atoms. In this way all of the atoms have a noble gas structure and they are electrically neutral.

VII. Isotopes

All atoms of a particular element contain the same number of protons (i.e. atomic number), but they may have different numbers of neutrons. Atoms exist as **isotopes**. Isotopes have the same number of protons, but a different number of neutrons. If the number of protons is changed a new element is formed. If the number of neutrons change, one merely has a new isotope of the **same** element.

carbon 12	6 protons	6 neutrons
carbon 14	6 protons	8 neutrons
nitrogen 14	7 protons	7 neutrons

The atomic mass or atomic weight of an atom represents the number of protons plus the number of neutrons in the nucleus. If you look at a periodic table, however, you will notice that atomic masses are rarely whole numbers. The atomic mass of an element given on the periodic table represents the **average relative mass of all naturally occurring isotopes** of that element.

The number of neutrons present is only important in that the higher the number of neutrons, the greater the mass of the isotope. However, **the chemical properties of an element depend solely on the number of protons**. Carbon 12 is **chemically** identical to carbon 14 or carbon 13. The reason that chemical properties depend on the number of protons is that the positively charged protons hold the negatively charged electrons in shells around the nucleus. Further, the exchange or sharing of electrons between atoms are the cause of chemical bonds.

VIII. Assignment (Part 3)

Part A: Fill in the blanks with the appropriate word or phrase:

1. Atoms with the same atomic number but a different atomic mass are called _____.
2. The results of the gold foil experiment led _____ to suggest that atoms are mostly empty space, but do contain a "solid" core which is called the _____.
3. Elements in group 15 have _____ electron(s) in the outer-most energy level.
4. The number of valence electrons equals the number of electrons in the _____ energy level.
5. The halogens have _____ valence electron(s).
6. Noble gases do not react at SATP because they have _____.
7. An atom that has lost or gained electrons is called a(n) _____.
8. If an atom gains electrons, it forms _____ charged ions called _____.
9. When electrons are lost from an atom it forms _____ charged ions called _____.
10. Elements that do not chemically react with other elements are said to be _____. An example is _____.
11. The alkali metals form ions with a _____ charge.
12. When elements in group 2 react with other elements, they _____ (lose/gain/share) _____ electrons.
13. When elements in group 16 react with other elements, they _____ (lose/gain/share) _____ electrons.
14. Metals tend to _____ (lose/gain/share) electrons, whereas non-metals tend to _____ (lose/gain/share) electrons.
15. When there is a transfer of electrons from one atom to another, a(n) _____ bond is formed.
16. When two atoms share electrons a _____ compound forms.
17. _____ compounds are formed when a metal reacts with a non-metal.
18. The subatomic particle that is much smaller than the others is the _____.
19. When two non-metals are combined they form a(n) _____ compound.
20. When a metal and a non-metal combine they form a(n) _____ compound.
21. The noble gas with electrons only in the first energy level. _____
22. The halogen that forms ions containing 18 electrons. _____
23. The element in period 3 containing 3 valence electrons. _____

Part B: Complete the chart below.

Atom or ion name	Symbol	Atomic mass	Atomic number	# of neutrons	# of protons	# of electrons	# of electrons lost/gained	Net charge
iron (III)	Fe^{3+}	56	26	30	26	23	lost 3	3+
					25	21		
sodium						10		
						10		3+
		40		22			0	
	F^-							
hydride								1-
			16				gained 2	
	Mg^{2+}							
						18	lost 2	
			1			0		
sulfur								
		59		32				3+

Part C: Complete the following table. Note the name of a *non-metallic* ion ends in *ide* while the name for a *metallic* ion uses the full name of the metal.

Ion name	Symbol	# of p ⁺	# of e ⁻	# of electrons lost/gained	Same electron structure as what Noble gas?
e.g. fluoride	F ⁻	9	10	gained 1	neon
1.		53	54		
2.		16		gained 2	
3. potassium				lost 1	
4.	Ca ²⁺				
5.		35	36		
6.	Sr ²⁺				
7.	H ⁺				(none)
8.		8		gained 2	
9.		12		lost 2	
10. aluminum			10		
11.		34	36		
12.	H ⁻				
13. lithium				lost 1	
14.	Rb ⁺				
15.		17	18		