Basic Chemistry Review

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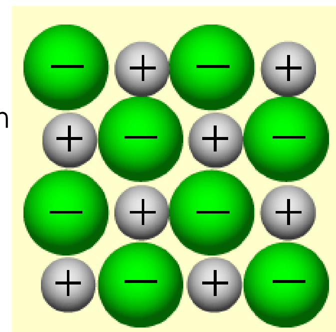
1. Review: J.J. Thomson's Plum Pudding model and ionic vs. molecular bonding

In Unit 6, we examined the interactions between positively charged objects (top tape), negatively charged objects (bottom tape), and neutral objects (foil and paper strips). To explain that objects can become charged, J.J. Thomson proposed that atoms have smaller mobile particles in them. Evidence from his Cathode Ray experiments showed that these smaller mobile particles, later called electrons, are negatively charged. For an atom to be neutral, there must be the same number of positive charges to counter balance the negative charges of electrons. Thomson had no experimental evidence to show where in the atom the positive charges would be. He hypothesized that the interior of an atom was a positive cloud with no mass. It is the attraction between the positive cloud and the negative electrons that holds electrons inside the atom.

However, atoms of different elements have different abilities to attract electrons. We call this electronegativity. Metal elements have free moving electrons that make metals good conductors of electricity. This suggests that the positive charges in metal atoms attract electrons weakly (low electronegativity). On the other hand, nonmetal elements are poor conductors, suggesting strong attraction (high electronegativity) between the positive charges and electrons. This limits the movement of the electrons between atoms. Because of this difference in electronegativity, a metal atom is more likely to form a positive ion (cation) when electrons are transferred to a nonmetal atom due to the higher attraction (higher electronegativity) from the positive charges in the nonmetal atom. The nonmetal atom becomes a negative ion (anion). Thus, the bonding between metal atoms and nonmetal atoms is ionic, as shown by the diagram to the right. If the difference in electronegativity between atoms is not enough to cause electrons to move from one atom to another, i.e., two nonmetal atoms such as H and Cl, these atoms bond together to form neutral HCl molecules, as illustrated in the diagram on the left. However, the slight difference in electronegativity between H and Cl causes uneven distribution of electrons within the molecule, with the H end of the molecule partially positive (δ+) and the Cl end partially negative (δ-). Thus we call HCl a polar molecule.



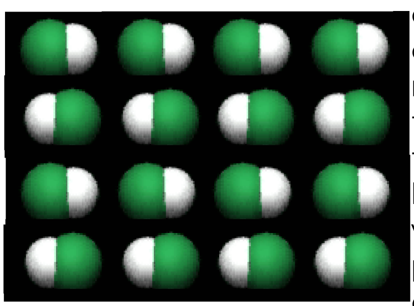


Table 6.2

Electronegativity Values for Selected Elements

H 2.1						
Li 1.0	Be 1.5	B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.7	Ba 0.9	Tl 1.8	Pb 1.9	Bi 1.9		

(Recall the tug-of-war analogy we used in describing this situation.)

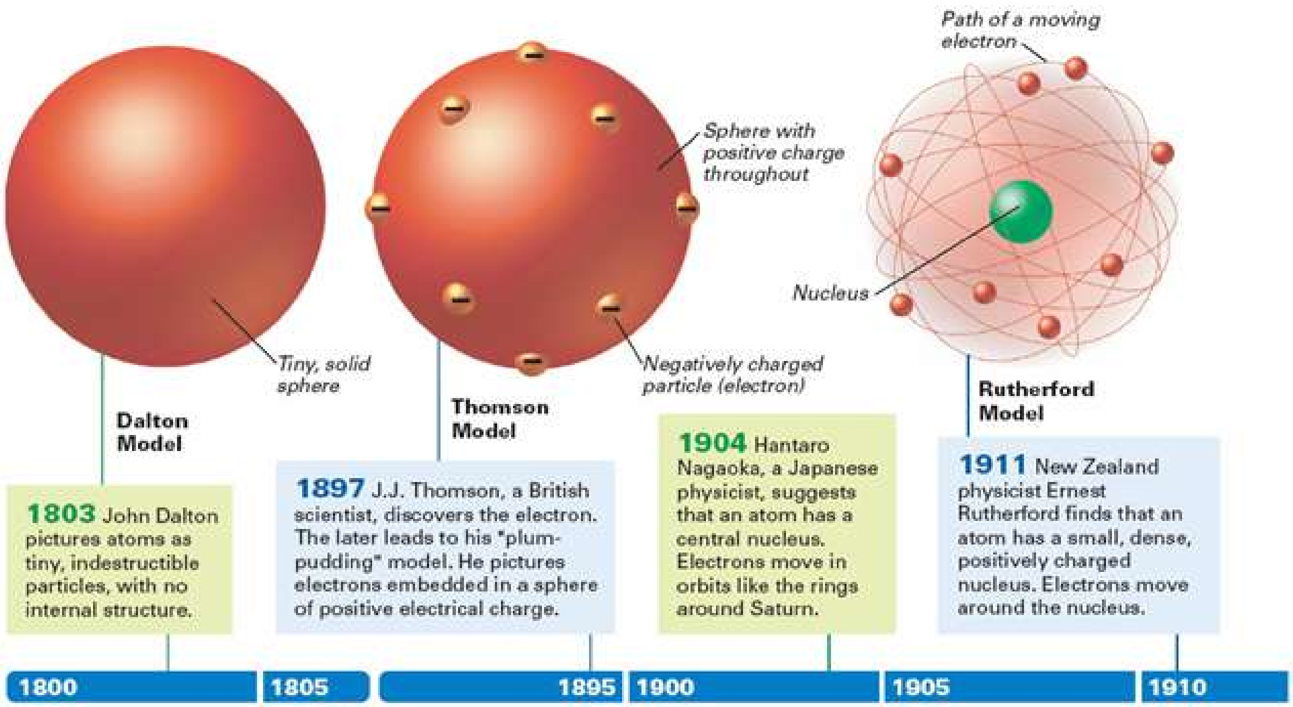
A numerical scale of 0 – 4 has been used to compare the electronegativity of main group elements, as shown in the table to the right.

Intermolecular forces: For polar molecules, the δ+ end of one molecule weakly attracts the δ- end of another molecule. This weak attraction between polar molecules is call dipole-dipole interaction – one of the intermolecular forces that plays vital role in biological systems.

2. Atomic models beyond Thomson's Plum Pudding Model

Over ten years after J.J. Thomson proposed his Plum Pudding model of an atom, Ernest Rutherford, a former student of Thomson's, proposed a nuclear model of an atom based on evidence collected from his famous gold foil experiment. (Check out the detailed explanation of his experiment at the following link. http://www.mhhe.com/physsci/chemistry/essentialchemistry/flash/ruther14.swf) Rutherford suggested that the positive charges in an atom are concentrated in a very small but dense center of the atom that he called a nucleus. Almost all the mass of the atom is also in the nucleus. Without any evidence on where electrons are in the atom, Rutherford hypothesized that electrons are moving around the nucleus. Thus, according to Rutherford's model, atoms are made of mostly empty space. Later, based on the experimental work by Henry Moseley, James Chadwick and others, scientists proposed that the nucleus is made of positively charged protons and neutral neutrons. Both protons and neutrons have mass, with neutrons slightly more massive than protons. Therefore the number of protons in the nucleus must be the same as the number of electrons so that the atom is neutral.

A comparison of Dalton, Thomson, and Rutherford model is shown in the diagram below.

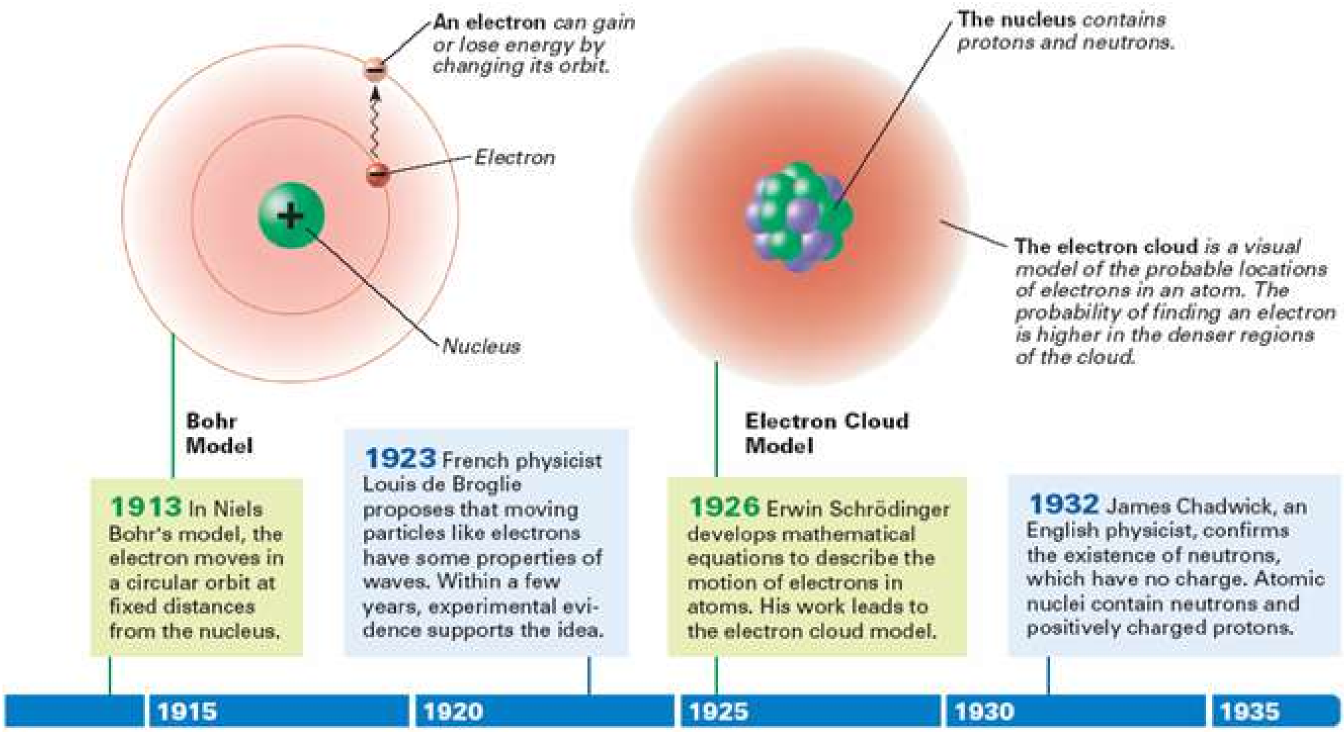


Meanwhile, other scientists, including Niels Bohr, Louis de Broglie, Werner Heisenberg, Erwin Schrodinger and others, took interests in studying the electrons of an atom. Bohr took the idea of quantized (discreet packets, or amounts) energy from Max Plank and Albert Einstein and calculated the energy levels for the electrons outside the nucleus. He proposed a model of an atom with electrons moving in circular orbits around the nucleus like the planets orbit the sun. Electrons in orbits closer to the nucleus are more attracted by the positive protons in the nucleus, therefore have less energy. (Recall in Unit 3 and Unit 7, we generalized that the more attracted the particles are to each other, the closer the particles are, therefore the less potential energy is stored.) The farther away from the nucleus, the more energy electrons have. However, his model could only explain the experimental results of hydrogen atoms which has only one electron in each atom. All the other elements have multiple electrons in their atoms. So a better atomic model was needed to explain experimental observations of all atoms.

Eventually, based on the work from many scientists, a modern quantum atomic model emerged. In this model, electrons do move around the nucleus with protons and neutrons in it. However, electrons do not have fixed orbits. Instead, electrons behave not only as discrete particles, but also as waves. (Recall in physics, we talked about light and other electromagnetic radiations as waves.) This is kind of hard for us to imagine. A cartoon video at the following link explains the basic idea of the dual property of electrons. http://www.youtube.com/watch?v=x_tNzeouHC4

Because of this particle-wave property of electrons, the actual location of an electron at any given moment cannot be determined. Based on the energy of an electron, we can only know the probability of electron appearance in the space outside the nucleus. Generally speaking, low energy electrons are mostly likely to appear in the space close to the nucleus while high energy electrons are more likely to appear in the space farther away from the nucleus. The outermost electrons of an atom have highest energy and are most "active". We call these electrons **valence electrons**. These electrons are the ones primarily involved in chemical bonding, which will be discussed in the next section.

The comparison of Bohr model and the quantum model of a single-electron atom is shown in the figure below. In the quantum model, the map of the probable locations of the electron is usually called electron cloud.

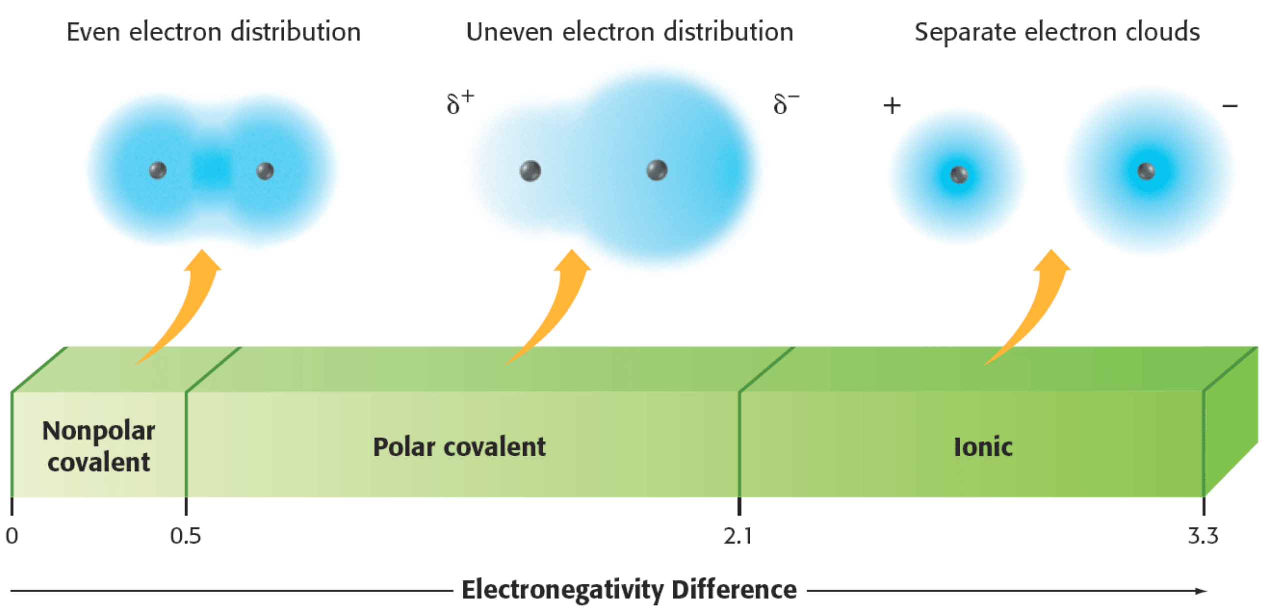

An electron can gain or lose energy by changing its orbit.
Electron
Nucleus
Bohr Model
The nucleus contains protons and neutrons.
The electron cloud is a visual model of the probable locations of electrons in an atom. The probability of finding an electron is higher in the denser regions of the cloud.
Electron Cloud Model
1913 In Niels Bohr's model, the electron moves in a circular orbit at fixed distances from the nucleus.
1923 French physicist Louis de Broglie proposes that moving particles like electrons have some properties of waves. Within a few years, experimental evidence supports the idea.
1926 Erwin Schrödinger develops mathematical equations to describe the motion of electrons in atoms. His work leads to the electron cloud model.
1932 James Chadwick, an English physicist, confirms the existence of neutrons, which have no charge. Atomic nuclei contain neutrons and positively charged protons.
1915
1920
1925
1930
1935

3. An evolved view on bonding based on the quantum atomic model

As discussed in the previous section, high energy valence electrons are farther away from the nucleus. They are less attracted by the positive nucleus. Further more, they are also repelled by the inner electrons, those electrons closer to the nucleus. As a result, valence electrons are more likely to "jump" to another atom when the nucleus of the other atom exert much more attraction to these electrons than their own nucleus. This is the case of ionic bonding between metal atoms and nonmetal atoms, as was discussed in the Thomson model. For example, in the compound NaCl, the significant difference in electronegativity between Na (0.9) and Cl (3.0) causes the transfer of the only valence electron of the Na atom to the Cl atom, thus forming Na^+ and Cl^- in this ionic compound.

In the case of molecular compounds, which are made of all nonmetal atoms, valence electrons are not able to completely "jump" from one atom to another because both nonmental elements have high electronegativity with not too much of a difference. Let's still use H and Cl as an example. The electronegativity of Cl is 3.0, while for H it is 2.1. The valence electron of H and one of the valence electrons of Cl are mostly found in the region between the two nuclei, in other words, the electron clouds overlap. This kind of electron sharing between nonmetal atoms is called covalent bonding. Since Cl has higher electronegativity, the most probable locations of finding these two valence electrons are closer to Cl than H, thus a polar covalent bond.

The following diagram illustrates the difference between nonpolar covalent, polar covalent and ionic bonds based on the difference in electronegativity between the two atoms.



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1) Look up Lewis Dot diagrams for H2 and F2 how are they different.

2) What, if anything, about these structures could account for the differences in their boiling points?

Element	boiling pt. (K)
Hydrogen	20
Nitrogen	77
Oxygen	90
Fluorine	85
Chlorine	239
Bromine	332
Iodine	457

3) Go to the following website: http://phet.colorado.edu/en/simulation/molecule-polarity and click on Run Now!

After the window loads, you should see a blue screen with two atoms A and B. They are connected together by a covalent bond.
On the right hand side of the screen, there will be a box that says "Surface." Click the check box that says "Electron Density." Two grayish spheres should show up with a key showing electron density at the bottom.

Using the electron density diagrams for these elements, explain the origin of the attractive forces that exist between the molecules. Why should they be greater for F2 than for H2?

4) Plot a graph of boiling pt. vs. number of electrons for these diatomic elements. What general relationship appears to exist between b.p. and number of electrons? Is this consistent with your answer to #3?

5) Charged objects exert forces on one another. Describe the distinction chemists make between *attractions* (discussed in #3) and chemical *bonds* (discussed in the last unit).

6) Examine the electron density map for methane CH4. Why is this molecule described as non-polar?

7) Describe the trend between boiling pt and number of electrons for the class of hydrocarbons called alkanes, C_nH_{2n+2}. Account for this trend in terms of your answers to questions 2 and 6.

alkane	boiling pt. (K)
CH4	111
C2H6	187
C3H8	231
C4H10	272
C5H12	309
C6H14	342

Now click on the tab that says “Real Molecules.” There are a number of molecules that we will continually refer to in biology. Out of that list oxygen, carbon dioxide and water.

8) Click on each three and look at them. Which have bond dipoles? Which have molecular dipoles? Which do you think are polar/nonpolar? Why are they polar/nonpolar?

9) If I had two water molecules near each other do you think they would interact? Would there be any intermolecular forces? If yes why? Draw what this interaction may look like.

Use the PhET simulation molecular-polarity to help you answer the following
Be sure to refer to diagrams to support your explanations.
10) Under the Real Molecules tab, view the electron density diagram for HF. How does it differ from the one for F2?

11) Examine the diagrams for HF and H2O. Explain why the distribution of electron density for molecules like HF and H2O is not symmetrical. What is meant by the term “electronegativity”?

12) Use the electrostatic potential diagrams for HF and H2O to explain why these molecules are said to be polar. How does this representation help explain the stronger attractions between these molecules?

13) How does the polar nature of water explain how ionic substances dissolve in water?
The Flash animation found at this site may help.
http://www.mhhe.com/physsci/chemistry/essentialchemistry/flash/molvie1.swf

14) Were it not for hydrogen-bonding, in what phase would water exist on our planet? Explain.

15) How do hydrogen bonds differ from dipole-dipole attractions? Why are hydrogen bonds called *bonds*, rather than attractions?

16) Describe the role hydrogen bonding plays in the structure of ice. Why is ice less dense than liquid water? Draw water molecules in the liquid phase and then in the solid phase to support your explanation.

17) Transpiration is the name given to how water travels up a plant, from its roots to its leaves. Discuss how this happens in the context of waters special properties.

18) Both ethanol, CH3CH2OH and dimethyl ether, CH3OCH3 have the same molecular formula, but one of these substances has a much higher b.p. than the other. Predict which has the higher b.p. and explain. Look up and compare the diagrams of each molecule to support your explanation.

19) What is pH? What does the pH scale represent? How do you calculate pH?

20) Why is pH important to the functioning of biological systems?