Bohr Model of the Atom:

* Named for the Danish physicist Neils Bohr (1885-1962) who developed the model.
* Bohr developed the model to explain the unique pattern of light that each element gives off when heated or electrified.
* Bohr’s model also explains many of the chemical properties of the elements that we will be studying.
* Like the planetary model, the protons and neutrons are in the nucleus and are responsible for (nearly) all of the mass of the atom. The electrons orbit the nucleus.
* What the Bohr model added was a distinct and strict pattern to the electron arrangement:
  + The electrons are arranged in **shells**, also called orbitals or energy levels.
    - The first (innermost) shell can hold 2 electrons
    - The second shell holds 8 electrons
    - The third shell holds 8 electrons
    - the fourth shell holds 18 electrons (2,8,8)
  + The INNER SHELLS are lower energy and FILL FIRST.
  + Single electrons fill before pairs.
* The number of electrons is determined by looking at the number of protons and the *ion charge*. Most atoms are electrically neutral.
* The outer most shell is called the *valence shell*. The electrons in the valence shell are called *valence electrons*.
* The valence electron arrangement largely determines the chemical behavior of the atom.

Examples:

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Ions:

* An ion is an atom that has *different numbers of PROTONS and ELECTRONS.*
* In other words: An ion is an atom that has *a NET ELECTRIC CHARGE.*

As we already know protons carry a charge of **+1** while electrons carry a charge of **-1**. Most atoms, and therefore most objects are *neutral* because they have an exact balance of protons and electrons.

However, an atom can gain or lose electrons as it randomly zips about, and when it does it ends up with an imbalance between electrons and protons. When this happens the atom will have a *NET or OVERALL CHARGE.*

* **IF AN ATOM HAS MORE PROTONS THAN ELECTRONS (IT HAS LOST ELECTRONS) IT WILL HAVE A NET \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ CHARGE AND WILL BE CALLED A \_\_\_\_\_\_\_\_\_\_\_\_\_\_ .**
* **IF AN ATOM HAS MORE ELECTRONS THAN PROTONS (IT HAS GAINED ELECTRONS) IT WILL HAVE A NET \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ CHARGE AND WILL BE CALLED AN \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ .**

As atoms fly around (very quickly! a molecule of O2 in the air at 20oC is moving at 2500km/h, on average) they can lose or gain electrons in all sorts of ways. However *atoms are most stable when then have a FULL VALENCE SHELL*. Atoms can get a full valence by gaining electrons (How many electrons would fluorine need to gain?) or by losing electrons (How many electrons would calcium need to lose?).

EXAMPLES: Draw a BOHR model of OXYGEN:

How many electrons must it LOSE or GAIN to have a full valence shell?

What is its ion charge?

Is this an ANION or CATION?

EXAMPLES: Draw a BOHR model of BERYLIUM:

How many electrons must it LOSE or GAIN to have a full valence shell?

What is its ion charge?

Is this an ANION or CATION?

EXAMPLES: Draw a BOHR model of CHLORINE:

How many electrons must it LOSE or GAIN to have a full valence shell?

What is its ion charge?

Is this an ANION or CATION?

EXAMPLES: Draw a BOHR model of SODIUM:

How many electrons must it LOSE or GAIN to have a full valence shell?

What is its ion charge?

Is this an ANION or CATION?

Science 10: Compounds

Compounds are formed when two or more elements combine. This can happen in two main ways.

1. **Ionic Compounds** are formed when a metal donates electrons to a non-metal. In the process the metal becomes a **cation** (+ charge) and the non-metal becomes an **anion** (- charge). The two (or more) ions then become stuck together (positives and negatives are attracted!) to form an ionic compound.

* Ionic bonds are always between a METAL and a NON-METAL

2. **Covalent Compounds** are formed when two (or more) non-metals share electrons. This type of bond is generally weaker than an ionic bond. The result is still that each atom ends up with full valence shell.

* Covalent bonds are always between NON-METALS

**Ionic Compounds:**

Example: Sodium + Chlorine

Extra electron

BOHR

Sodium Chlorine

Empty Spot

LEWIS DOT

* Sodium has one extra electron, Chlorine has one too few; It’s a match made in heaven!
* So sodium can donate one electron to chlorine and both will end up with a full valence shell.
* To do this the atoms must move close together.
* When sodium gives up one electron it becomes an ion with a charge of 1+. When chlorine accepts an extra electron it becomes an ion with a charge of 1- (**A chloride ion**).
* These two ions then ‘stick’ with a force of *electrostatic attraction* together forming a NaCl molecule. The **net charge** of the molecule will be 0.

LEWIS DOT BOHR

Na Cl

The positive sodium ion and the negative

chloride ion are now electrically attracted.

Na+ Cl-

+ - The two become ‘stuck’ together with an *electrostatic force*. This is a new molecule!

NaCl

The last diagram(s) show a single molecule of NaCl, Sodium Chloride. Notice that the net charge of the MOLECULE is ZERO. The formula and name are always written with the metal element first.

Ex 2: Magnesium and Fluorine.

Florine Magnesium

* This is a little different. Magnesium has two extra electrons, fluorine only need one.
* In order for the compound to form **all atoms must end up with full valence shells**.
* This means we will need two fluorine atoms.
* Magnesium can donate both electrons, each fluorine can accept one.
* The atoms must move close together.
* Mg becomes Mg2+. Each F becomes F- (fluoride).
* The ions then ‘stick’ together with a force of *electrostatic attraction* forming a MgF2 molecule. The **net charge** of the molecule will be 0.

LEWIS BOHR

When the electrons leave the outer (3rd)

shell of Mg, that shell is empty, and so it

is NOT A SHELL AT ALL!

F Mg F

Notice that the atoms all *look* the same.

What makes them different?

F- Mg2+ F-

- 2+ -

Now each atom has a FULL VALENCE

SHELL.

MgF2

The last diagram shows a single molecule of Magnesium Fluoride.

**RULES FOR NAMING IONIC COMPOUNDS:**

1. Write the name of the **metal element** first.

b. If the metal is **multi-valent** include a roman numeral to indicate ion charge.

2. Write the name of the poly atomic ion or non-metal element next. For non-metal elements change the ending to **“ide”**

**Your turn:**

1. Beryllium + oxygen. For this example draw all three steps showing, with arrows how the electrons move. At the last step name the compound.

2. Sodium + Oxygen. For this example draw all three steps showing, with arrows how the electrons move. At the last step name the compound.

3. Lithium + Nitrogen. For this example see if you can skip right to the final molecule. Name the compound.

4. Magnesium and Nitrogen. For this example see if you can skip right to the final molecule. Name the compound.

5. Beryllium and fluorine. Try to skip the diagram. Just write the formula and name.

6. Calcium and phosphorous.

7. Aluminum and oxygen.

8. Aluminum and nitrogen.

**Covalent Compounds**

**Covalent Compounds** are formed when two (or more) non-metals share electrons. This type of bond is generally weaker than an ionic bond. The result is still that each atom ends up with full valence shell. We will be looking primarily at *BINARY COVALENT COMPOUNDS* meaning only two non-metal elements involved.

* Covalent bonds are always between NON-METALS
* Covalent bonds DO NOT INVOLVE a transfer of electrons. Instead atoms SHARE VALENCE ELECTRONS
* Unlike ionic bonds, two elements can form covalent compounds in different proportions. Examples include H2O (water) and H2O2 (dihydrogen dioxide, hydrogen peroxide), CO (carbon monoxide) and CO2 (carbon dioxide), CCl4 (carbon tetrachloride) and C2Cl6 (dicarbon hexachloride).

Example Fluorine + Chlorine

BOHR:

Fluorine Chlorine

LEWIS:

* Fluorine has one too few electrons, so does chlorine. How can both be happy?
* They can share. It is easiest to see this with a diagram. To make it clearer the valence shell is drawn extra-large and the fluorine has been turned upside down.

BOHR:

F Cl

LEWIS:

* The atoms must move close together, until their valence shells cross.

BOHR:

F Cl

LEWIS:

* Oh my! Take a second and count the electrons around the outer shell of fluorine (1,2,3…8!). Now count the electrons around chlorine (1,2,3…8!). HOORAH!

This is an example of a covalent compound. Valence electrons are shared.

**Oxygen and Bromine**

**Water and Hydrogen Peroxide: (H2O and H2O2)**

**Carbon and Hydrogen: (CH4 and C2H6)**

**Carbon Dioxide (double bonds)**

**Naming Covalent Compounds**

**F Cl**

The name of this molecule is chlorine monofluoride. The naming of *binary* (two elements) covalent compounds is a little bit tougher than for ionic compounds. First we need to know some prefixes.

|  |  |  |  |
| --- | --- | --- | --- |
| **Number** | Prefix | **Number** | Prefix |
| 1 |  | 6 |  |
| 2 |  | 7 |  |
| 3 |  | 8 |  |
| 4 |  | 9 |  |
| 5 |  | 10 |  |

|  |
| --- |
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| **Rule 1.** The element with the lower group (family) number is written first in the name; the element with the higher group (family) number is written second in the name.  ***Exception: when the compound contains oxygen and a halogen, the name of the halogen is the first word in the name.***  **Rule 2.** If both elements are in the same group, the element with the higher period number is written first in the name.  **Rule 3.** The second element in the name is named as if it were an anion, i.e., by adding the suffix *-ide* to the name of the element.  **Rule 4.** Greek prefixes (see the Table provided at the top of this page) are used to indicate the number of atoms of each nonmetal element in the chemical formula for the compound.  ***Exception: if the compound contains one atom of the element that is written first in the name, the prefix "mono-" is not used.***  *Note: when the addition of the Greek prefix places two vowels adjacent to one another, the "a" (or the"o") at the end of the Greek prefix is usually dropped; e.g., "nonaoxide" would be written as "nonoxide", and "monooxide" would be written as "monoxide". The "i" at the end of the prefixes "di-" and "tri-" are never dropped.* |