

# Why and how atoms combine

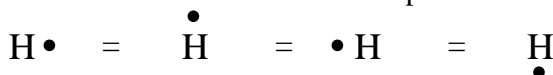
## Lewis diagrams

When atoms combine, only electrons in the outer (valence) shell are involved. We can represent these valence electrons with Lewis diagrams. Lewis diagrams for the first 20 elements are shown below.

IA	IIA	IIIA	IVA	VA	VIA	VIIA	0
H•							•He•
Li•	•Be•	•B•	•C•	•N•	•O•	•F•	•Ne•
Na•	•Mg•	•Al•	•Si•	•P•	•S•	•Cl•	•Ar•
K•	•Ca•						

To write Lewis diagrams:

1. Write the element symbol. Around this draw dots – one for each valence electron.
2. The dots should be spread over four sides. Dots are not paired until all sides have at least one dot.
3. It does not matter on which side dots are placed. For example, hydrogen can be drawn four ways:



4. The number of valence electrons is equal to the group number. For example, hydrogen is in group IA (group 1) and it has one valence electron. Neon is in 0 (group 8) and it has 8 valence electrons. The only exception is He which is in group 8 but has 2 valence electrons.

Q1 – Write Lewis diagrams for a) Ne, b) Sb (Z=51), c) Rb, d) a neutral atom with 9 total electrons

Q2 – Write all possible variations of the Lewis diagram for phosphorus.

## The Octet Rule

The noble gasses do not react with other elements. In other words, noble gasses are very stable. Let's look at the electron configuration of noble gasses (note: K is the name given to shell 1, L to 2, etc.)

Element	Shell					
	K	L	M	N	O	P
He	(2)					
Ne	2	(8)				
Ar	2	8	(8)			
Kr	2	8	18	(8)		
Xe	2	8	18	18	(8)	
Rn	2	8	18	32	18	(8)

Most Noble gasses have 8 electrons in their outer shell. This is a stable conformation.

### The Octet Rule

When atoms form ions or combine in compounds they obtain electron configurations of the nearest noble gas (usually this means that there will be 8 outer electrons).

## The Octet rule and ion formation

As stated, when atoms form ions they seek an electron configuration like that of the nearest noble gas. Nearest refers to the number of representative elements between an element and a noble gas. For example Br is 1 space removed from Kr, but 7 spaces removed from Ar (see your periodic table). So, Br would gain 1 electron to form Br<sup>-</sup>. Similarly, Mg would lose 2 electrons to become Mg<sup>2+</sup>.

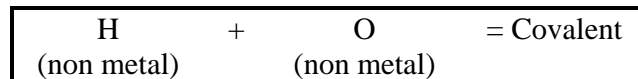
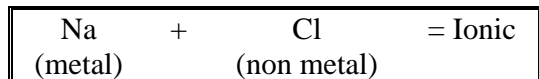
Q3–Complete the chart:

	Br	P	Ne	Al	Ca
Nearest Noble gas (spaces removed)	Kr (1)				
Resulting ion	Br <sup>-</sup>				

Group IVA elements are equal distances from two noble gasses. These elements tend not to form ions.

## Types of bonds

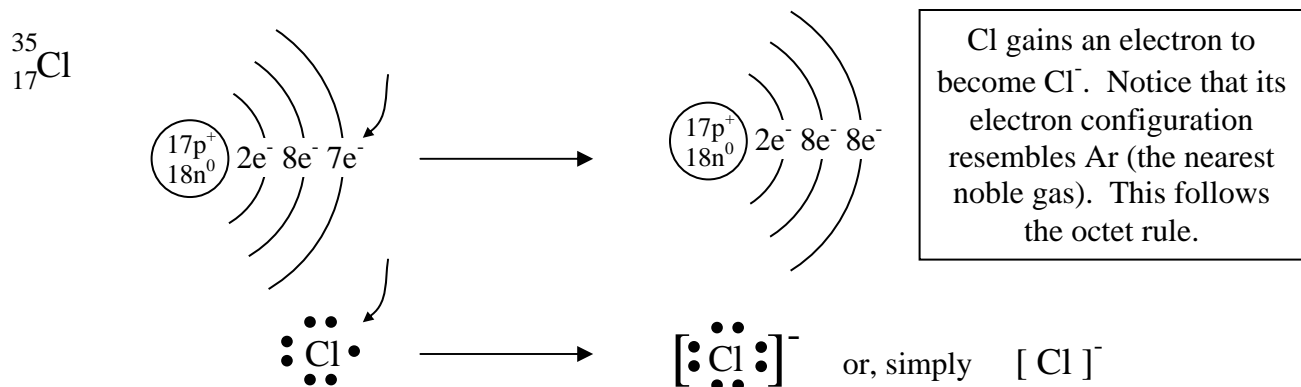
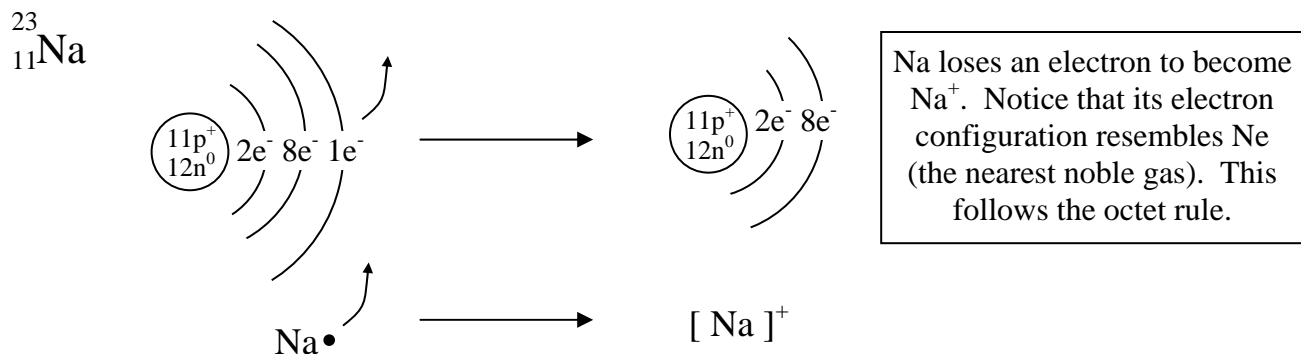
There are two types of bonds: ionic and covalent. Ionic bonding occurs between metals and non-metals. Covalent bonding occurs between two non-metals. (When two metals combine an alloy is formed. Alloys are unusual substances; we will not concern ourselves with bonding in alloys)



Q4 – Identify each compound as ionic or covalent: MgO, CaCl<sub>2</sub>, SO<sub>2</sub>, PbCl<sub>2</sub>, CCl<sub>4</sub>, CH<sub>4</sub>

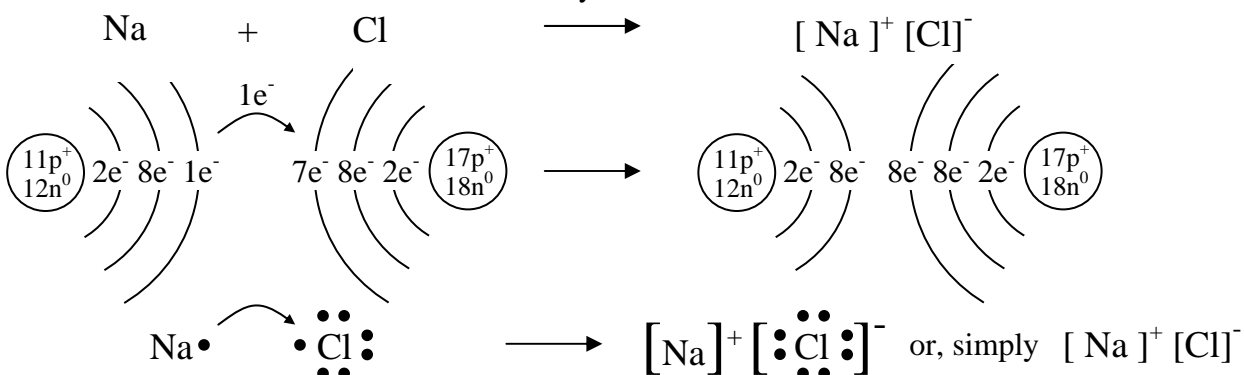
### Formation of ions

Ionic bonding involves the formation of ions: metals lose electrons to become positive ions, non-metals gain electrons to become negative ions. This can be shown with Bohr-Rutherford diagrams and Lewis diagrams:



### Ionic bonding

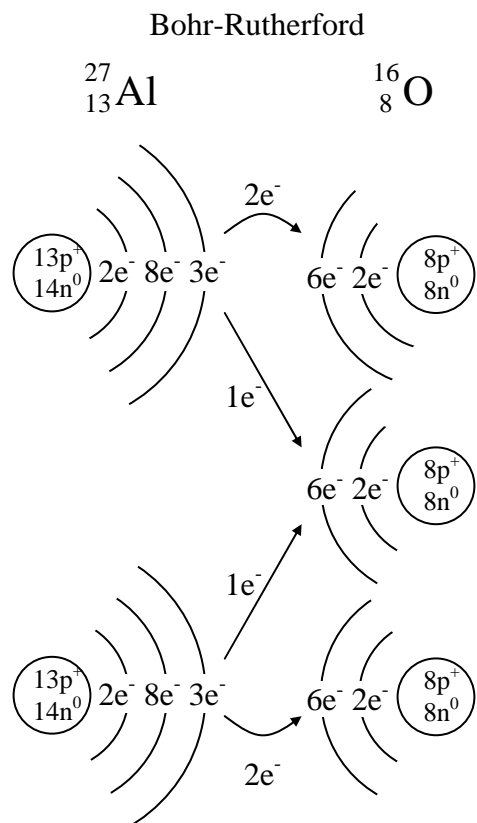
The formation of individual ions, as shown above, does not usually happen in nature. Instead, the formation of a positive ion is always coupled with the formation of a negative ion. For example, when Na combines with Cl, the Na becomes positive and the Cl becomes negative. The positive and negative ions are then attracted to each other. There are three ways to illustrate this:



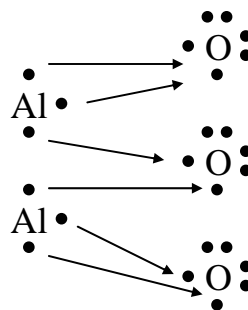
Q5 – Use all three methods (chemical reaction equation, B-R diagrams, and Lewis diagrams) to show how bonds form between: Li + Cl, Mg + O (follow the octet rule)

## Ionic bonding continued

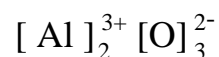
Sometimes ionic bonding is more complex than illustrated above. Sometimes it will be necessary to use more than one atom to balance electrons. Just remember to keep adding atoms until the octet rule is achieved. For example, let's look at O + Al,



Lewis diagrams



Either way we draw it, two aluminum ions are formed (each with a charge of 3+), and three oxygen ions are formed (each with a charge of 2-). This can be written as:



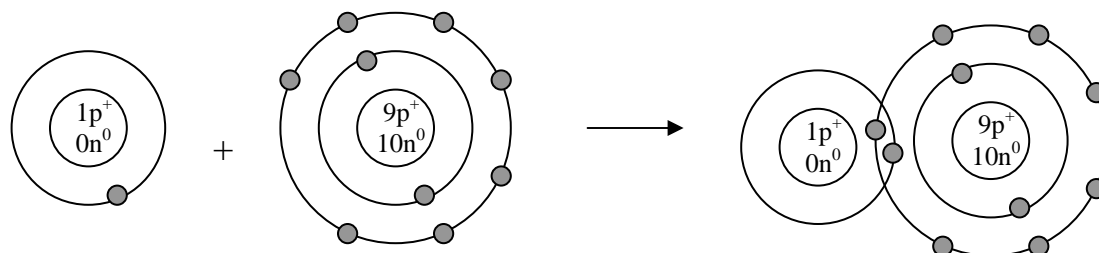
So, the chemical reaction equation would be  $2\text{Al} + 3\text{O} \rightarrow \text{Al}_2\text{O}_3$  (actually,  $2\text{Al} + 1\frac{1}{2}\text{O}_2 \rightarrow \text{Al}_2\text{O}_3$ ). Or if we want to indicate charges we could write:  $2\text{Al} + 3\text{O} \rightarrow [\text{Al}]_2^{3+} [\text{O}]_3^{2-}$

Q6 – Use all three methods (chemical reaction equation, B-R diagrams, and Lewis diagrams) to show how bonds form between: Li + O, Al + Cl (follow the octet rule)

## Covalent bonding

Covalent bonding occurs between two non-metals. Covalent bonding is different from ionic bonding because electrons are shared instead of transferred. Yet, covalent bonding still follows the octet rule. Let's look at the bond that is formed between H and F.

Bohr-rutherford



Lewis diagram

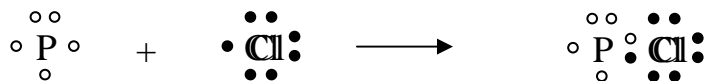


Notice that when H and F share electrons that each is able to complete its outer shell.

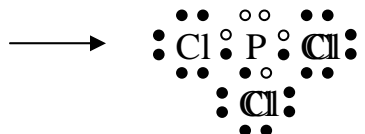
## Covalent bonding continued

Just like with ionic bonding, sometimes one atom of each element is not enough. In these cases, we will have to continue to add atoms until the octet rule is satisfied. Let's use Lewis diagrams to illustrate the reaction of P and Cl,

First, we know that this is a covalent bond because P and Cl are both non-metals.



This is a good start, but we have a problem: P does not have an octet. To give P an octet we must add two more atoms of Cl.



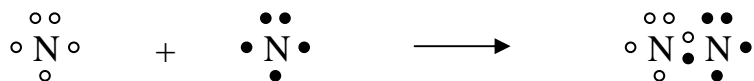
Now each atom has a filled outer shell. You should not be concerned that the number of Cl atoms is not equal on each side. Eventually we will concern ourselves with this. For now, think of the equation as a question: when P and Cl react, in what ratio do they combine?

Q7 – Draw Lewis dot diagrams to represent the following compounds (think about whether they are ionic or covalent):  $\text{CCl}_4$ ,  $\text{HCl}$ ,  $\text{MgF}_2$ ,  $\text{H}_2\text{O}$ ,  $\text{NH}_3$ ,  $\text{NaCl}$ ,  $\text{OH}^-$ ,  $\text{H}_2$

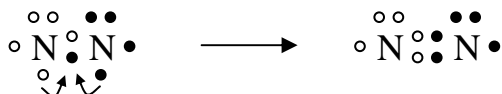
## Multiple bonds

Sometimes when atoms combine, there is no way to add them together to fill the outer shells of each atom. For example, if you add N and N, you can keep adding N, but you will not be able to fill the outer shell of all the atoms. In this case a different approach is required.

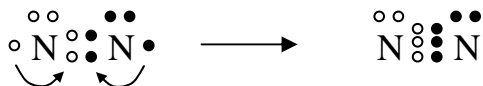
Let's start by combining N and N,



We are not done because each N has only 6 valence electrons. In this case we need to add another "bond" between the nitrogen atoms. To create another bond we move one unpaired electron from each nitrogen atom into the middle of the two atoms:



Still we do not have 8 electrons around each nitrogen, so we have to add another bond:



Now there are 8 electrons surrounding each nitrogen atom. Often multiple bonds are drawn with the electrons spread out:  $\begin{array}{c} \circ \circ \bullet \bullet \\ \circ \text{N} \bullet \bullet \text{N} \bullet \\ \circ \bullet \bullet \end{array}$  or with lines representing bonds:  $\begin{array}{c} \circ \circ \bullet \bullet \\ \circ \text{N} \equiv \text{N} \bullet \\ \circ \bullet \bullet \end{array}$

## Lewis structures

As you saw above with nitrogen, pairs of electrons can be represented with a line instead of two dots (one per atom).

Q8 – Draw Lewis dot diagrams and Lewis structures, where possible, for the following compounds:  
 $\text{HCl}$ ,  $\text{CO}_2$ ,  $\text{Na}_2\text{O}$ ,  $\text{NH}_3$ ,  $\text{O}_2$ ,  $\text{I}_2$ ,  $\text{Al}_2\text{O}_3$ ,  $\text{O}_3$ ,  $\text{CH}_4$