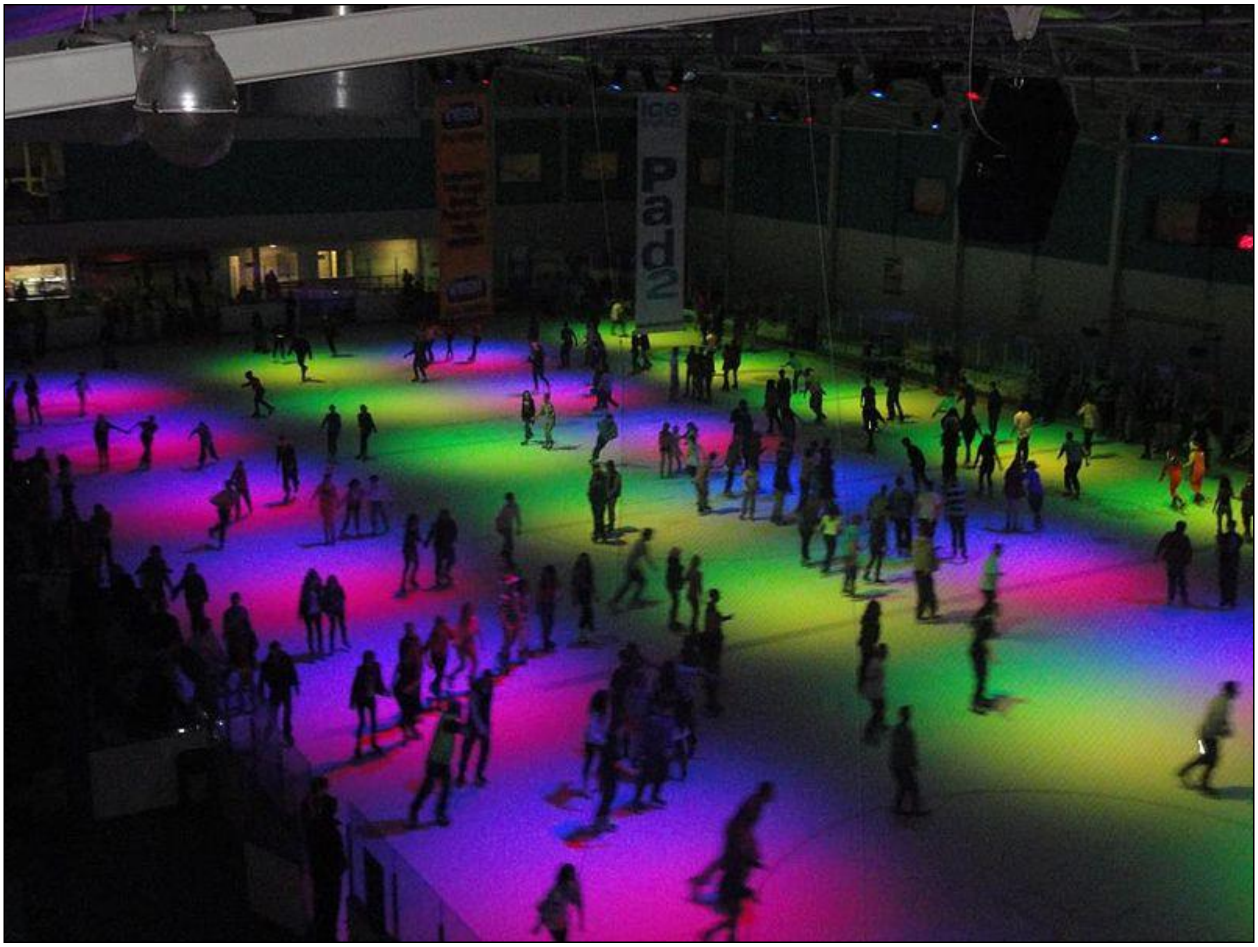
Chapter 8: Ionic Compounds



Ionic compounds are frequently used in competitive roller disco.

http://commons.wikimedia.org/wiki/File:Disco_sheff.jpg

Chapter 8: Ionic Compounds

Now, I want to start off this chapter by debunking a commonly-held myth: The myth of the ionic bond. *There is no such thing as an ionic bond!* By saying that ionic bonds exist, you are exposing yourself as a common and uncultured hoodlum. Woe to the one who uses the term "ionic bond!"[1]

Section 8.1: The Formation of Ionic Compounds

When a nonmetal and a metal react with one another, the result is an ionic compound. The term **ionic compound** refers to a compound where positively-charged **cations** are attracted to negatively-charged **anions** due to the attraction of their opposite charges. This attraction between different charges is referred to **electrostatic attraction** and this attraction is what's normally misnamed an "ionic bond."

So, how do these ions form in the first place? If you guessed that it has something to do with the octet rule[2], you're right. Let's see what happens when neutral fluorine and lithium atoms react to form the ionic compound lithium fluoride.

- By itself, fluorine has seven valence electrons. However, due to the octet rule (which states that all elements want to lose or gain electrons to be like the nearest noble gas), there is a strong driving force for it to gain one electron to get the same electron configuration as neon.

- Likewise, sodium has one valence electron. However, due to the octet rule, there is a strong driving force for it to lose this one electron to get the same electron configuration as helium.

- When a lithium atom comes into contact with a fluorine atom, the lithium atom (which wants to lose an electron[3]) gives its electron to the fluorine atom (which wants to gain an electron). This loss of an electron makes lithium into a Li^+ cation and the gain of an electron makes fluorine into a F^- anion.

- Because positive and negative charges are attracted to one another, the ions to combine to form the ionic compound LiF.

- More than two atoms typically combine with each other to form larger ionic crystals. The pattern in which the ions arrange themselves is called a **crystal lattice**. The smallest repeating unit that makes up the crystal lattice is called the **unit cell** of the compound. This can be imagined by thinking of a child playing with "Legggo" blocks.[4] Each block is equal to a unit cell, and the whole giant structure that's ultimately made is the whole crystal.

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[1] Seriously, don't use the term "ionic bond." It's not chemically-valid.

[2] See the last chapter.

[3] I commonly refer to atoms or molecules "wanting" to do things. The reader should understand that a more rigorous way of describing this would be to say that an atom is *more stable* when something happens.

[4] Extra "g"s added as the result of a lawsuit for copyright infringement.

Section 8.2: Properties of Ionic Compounds

It's probably not surprising that the properties of ionic compounds are closely related to the way in which cations and anions have been combined with one another. For your convenience, I have included the crystal structure of sodium fluoride as a reference in this section:

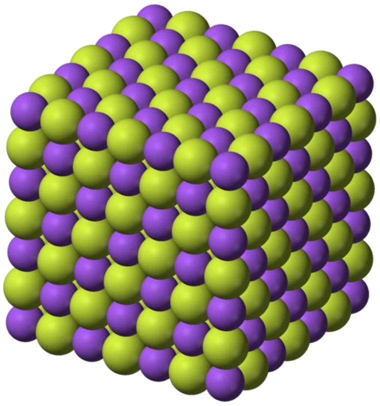

Figure 8.1: *The crystal structure of NaF. The strong electrostatic interactions between the cations and anions are responsible for the properties of ionic compounds.*

http://commons.wikimedia.org/wiki/File:Sodium-fluoride-3D-ionic.png

One term that's often used to describe how strongly the ions in an ionic compound are bound to each other is called the **lattice energy,** which is the amount of energy it takes to pull one mole of ions in a compound apart from one another.[5] The lattice energies of ionic compounds are usually highly negative, which means that there's a lot of energy needed to pull apart the ions.

Using this knowledge and Figure 8.1 as a guide, let's examine the properties of ionic compounds:

- **Ionic compounds are hard**. When you consider that you have a bunch of positive and negative charges all attracted to one another in a big block, it's not really surprising that they'd be hard to move out of place. This resistance to movement makes ionic compounds hard.

- **Ionic compounds are brittle**. The ions in an ionic compound are lined up in just the right way so that cations and anions are arranged next to each other. However, when you add force to these crystals, the ions are moved relative to one another and don't necessarily line up correctly anymore. The resulting instability causes the crystal to shatter.

- **Ionic compounds have high melting and boiling points**. Again, if you've got a bunch of positively- and negatively-charged ions attracted to one another in a big block, it will take a lot of energy for them to be pulled apart and melt the compound. Many ionic compounds require temperatures of many hundreds of degrees Celsius to melt.

- **Ionic compounds conduct electricity when melted or dissolved**. Moving ions is one way that electricity can be conducted and when you either melt or dissolve an ionic compound, ions are now free to move around. Chemical compounds that conduct electricity when dissolved are referred to as **electrolytes**.

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[5] We haven't discussed moles yet in any detail, so if this definition doesn't make much sense, don't worry about it.

Things With The Same Name:

Figure 8.2: *The top of this column is referred to as an "ionic capital" because this style of columnar cap originated in Ionia in the 6th century BCE.*

Photo courtesy of Ad Meskins.

Section 8.3: Ionic Naming and Formulizing

One of the best parts of being a chemist is using incomprehensible terms that nobody else understands. An excellent way to fulfill this goal is to learn how to name chemical compounds. After all, it's easy to say "table salt" when referring to a chemical compound, but it sounds a lot more impressive to call it "sodium chloride." In this section, we'll learn how to name ionic compounds, which is the first step to sounding smart.

Determining the Formulas of Ionic Compounds

Ionic compounds have two word names. The first word is the name of the cation and the second is the name of the cation. If you can get the names of these right, then you're in good shape.

- Record the formula and charge of the cation in the compound. Cations are usually metallic ions[6] and the charges of these ions can be determined either by their position on the periodic table[7], or the Roman numeral that's written in the name of the compound. For example: In the compound "sodium chloride", sodium has a charge of +1 because it needs to lose one electron to be like the nearest noble gas, and in "iron (III) sulfate" the iron has a +3 charge because there's a big Roman numeral "III" written in the name of the formula. Record the formula of the cation as well as its charge.

- Record the formula and charge of the anion in the compound. Anions are either nonmetallic ions or polyatomic ions. Polyatomic ions are ions that contain more than one atom, such as the sulfate ion (SO_4^{-2}). To find the charge of a nonmetallic anion, use the octet rule to determine how many electrons are needed to be like the nearest noble gas. For example, in the compound "sodium oxide", the oxide ion has a -2 charge because oxygen needs to gain two electrons to be like neon. As for polyatomic ions, you pretty much just need to memorize their formulas and charges. Table 8.1 shows you some of the most important ones that you should probably learn:

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[6] The main exception is the ammonium ion, NH_4^{+1}.

[7] The exceptions are zinc, cadmium, and silver. Zinc and cadmium always have a charge of +2, and silver always has a charge of +1. Unfortunately, you just have to memorize these.

***Table 8.1:** The common polyatomic ions and their charges. You should probably memorize these.*

Name of polyatomic ion	Formula and charge
acetate	$C_2H_3O_2^{-1}$ or CH_3COO^{-1}
ammonium	NH_4^+
bicarbonate	HCO_3^{-1}
carbonate	CO_3^{-2}
chromate	CrO_4^{-2}
cyanide	CN^{-1}
dichromate	$Cr_2O_7^{-2}$
hydroxide	OH^{-1}
nitrate	NO_3^{-1}
nitrite	NO_2^{-1}
permanganate	MnO_4^{-1}
phosphate	PO_4^{-3}
phosphite	PO_3^{-3}
sulfate	SO_4^{-2}
sulfite	SO_3^{-2}

- Write the formulas of the cation and anion next to one another. For example, in the compound "sodium sulfate" write the formulas of the sodium ion (Na^{+1}) and the sulfate ion (SO_4^{-2}) next to each other.

- In the final formula of the compound, the number of cations is equal to the charge on the anion, and the number of anions is equal to the charge on the anion. In the example above, there are two sodium ions in sodium sulfate because sulfate has a charge of -2. Likewise, there is one sulfate ion because sodium has a charge of +1. The formula, then, is written Na_2SO_4 to indicate the presence of two sodium ions and one sulfate ion.[8]

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[8] Some people mistakenly believe that the presence of "SO_4" in a compound formula means that there are four sulfate ions. This isn't true – the "4" is part of the name of the ion itself as shown in Table 8.1. It is only when a polyatomic ion has parentheses around it that we assume there is more than one of them, as is the case with $Ca(OH)_2$, which contains two hydroxide ions.

Determining the Names of Ionic Compounds

Another part of sounding smart is being able to write the names of ionic compounds. In this section, we will explore how to do this. To write the names of ionic compounds, follow these steps:

1. Write down the name of the cation. This is usually just the name of an element. For example, in FeF_2, the name of the cation is "iron."

2. Write down the name of the anion. This is either the name of an element with "-ide" added to the end, or the name of a polyatomic ion. In the case of FeF_2 above, the name of the anion is "fluoride."

3. Some compounds require Roman numerals between the name of the cation and anion. These Roman numerals correspond to the positive charge on the cation. Elements that require Roman numerals include the transition metals (except for Zn, Cd, and Ag) and the elements Sn, Pb, and Bi. If your cation is not one of the elements mentioned above, then *stop writing the name of the formula, because you are done.* In our case, iron is a transition metal, so we need a Roman numeral.[9]

4. If (and only if) your compound needs a Roman numeral, determine what this is with the following formula:

$$\text{Roman numeral} = \frac{\text{(charge on anion)(number of anions)}}{\text{number of cations}}$$

Given this formula, the Roman numeral for iron will be:

$$\text{Roman numeral} = \frac{(-1)(2)}{1} = 2 = \text{II}$$

And the name of the compound is "iron (II) fluoride."

Explanation of a Possibly Confusing Thing

The numbers in the equation above for FeF_2 came from the following locations:

- The charge on the anion is equal to -1, because fluorine is the anion and needs to gain one electron to be like the nearest noble gas.
- The number of anions is 2, because there is a "2" written under the formula of the anion.
- The number of cations is 1, because there is no other number written under Fe.

Multiplied together in the way shown, they give a Roman numeral of II.

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[9] Because the only time most people ever see Roman numerals anymore is when watching the Super Bowl, the first eight Roman numerals are I, II, III, IV, V, VI, VII, VIII.

Section 8.4: Metals

In every textbook, metals are described in the same chapter as the ionic compounds. Though there is, as far as I can tell, no reason for this to be the case, I will follow this convention to make learning easier for those already using other textbooks.

Metallic bonding occurs via something often referred to as the **electron sea theory**. In this theory, positively-charged metal nuclei are the islands in an ocean of electrons (Figure 8.3). These **delocalized electrons**[10] aren't stuck to any particular atom, but can move between the metal nuclei at will. Because the bonds formed by delocalized electrons can be easily rearranged, metals are malleable and ductile. Delocalized electrons also cause metals to be electrically-conductive. Because the bonding in a metal consists of mobile electrons, it's easy for applied electrons to travel through the metal as well.

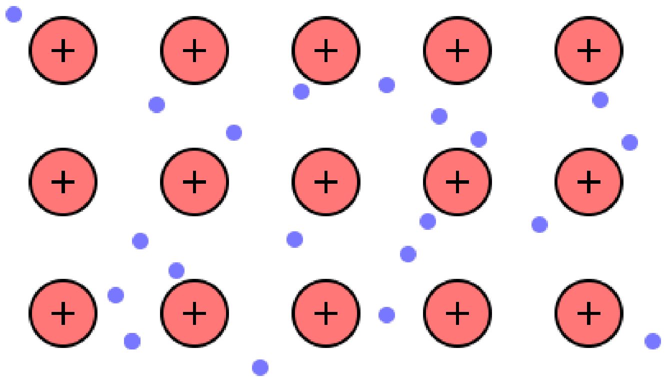

Figure 8.3: *In the electron sea theory, delocalized electrons in a metal travel freely between positively charged nuclei. This causes metallic bonds to be easily rearranged when force is applied to a metal.*

http://commons.wikimedia.org/wiki/File:Nuvola_di_e

Metals tend to have high melting and boiling points, due to the strength of metallic bonding. Metals are also shiny due to interactions of light with the delocalized surface electrons.

Spotlight on Technology Stuff

Figure 8.3: ***Alloys*** *are mixtures in which the main component is a metal. One sort of alloy, a* ***substitutional alloy****, occurs when the atoms of one element replace those of another. This occurs in brass, where some tin atoms replace copper atoms. In* ***interstitial alloys****, the atoms of one element fit between those of another. This occurs in steel, where carbon atoms fit between atoms of iron.*

http://en.wikipedia.org/wiki/File:Alloy_atomic_arrangements_showing_the_different_types.jpg

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[10] "Delocalized" refers to the fact that the electrons don't spend any length of time in any particular location. This is different than the "localized" electrons present in covalent bonds.

Chapter Summary

- Ain't no such thing as an ionic bond. Ions are held together with electrostatic attractions, *not* bonds.
- Ionic compounds occur when anions and cations are attracted to each other via electrostatic forces. The fact that they eventually form giant crystals of such attracted ions is responsible for many of the properties of ionic compounds.
- Metallic bonds occur when delocalized electrons surround a bunch of metal nuclei. These bonds aren't rigidly stuck in place, which helps to explain why metals are bendy.

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