Guided Practice Activities

**Module**: Bonding

**Section**: Ionic Compounds - **Key**
Ionic vs Covalent Compounds

**Activity 1**

The purpose of this activity is to recognize, name, and define different types of compounds, namely ionic and covalent compounds.

In general atoms of elements come together to form compounds. One can typically predict the type of compound based on the types of elements that have come together to form compounds. In this unit we focus on two types of compounds – ionic and covalent.

1. Ionic compounds are typically formed between metals and nonmetals.

2. The metal tends to lose electrons and form a cation with a positive charge. While the nonmetal in the compound tends to gain electrons and form an anion with a negative charge. The subsequently formed ions are held together in the compound by Coulombic attraction. Further the charge on the ion depends on how many electrons must be lost (in the case of a cation) or gained (in the case of an anion) in order for the electron configuration of that particular ion to be isoelectronic with the noble gas closest to it on the periodic table.

3. Covalent compounds are typically formed between nonmetals and nonmetals.

4. Chemical bonding involves valence electrons. In the context of ionic compounds the valence electrons in atoms helps to predict what? In the context of the covalent compounds the valence electrons are involved in what?

   In the context of ionic compounds, the number of valence electrons helps us predict the charge of the ions. For example, lithium has one valence electron and will give that electron up becoming a +1 cation and so achieve noble gas configuration. Calcium, however, has two valence electrons and will give up both of those electrons to become a +2 cation and so achieve noble gas configuration. In the case of the nonmetals, they seek to gain electrons. Oxygen has six valence electrons and will gain two more electrons becoming a −2 anion and so achieve noble gas configuration. Chlorine has seven valence electrons and will gain just one more electron becoming a −1 anion and so achieve noble gas configuration.

   In the context on covalent compounds, the valence electrons are involved in the shared bonds between atoms. Knowing the number of valence electrons for an element helps predict which atoms are bonded to each other.
ACTIVITY 2

A key skill in understanding both ionic and covalent bonding is the ability to predict the number of valence electrons in an atom.

1. Write out the electronic configuration of the following elements and state the number of valence electrons in each atom: Li, Be, B, C, N, O, F and Ne.

   - Li: $1s^22s^1$ – 1 valence electron
   - Be: $1s^22s^2$ – 2 valence electrons
   - B: $1s^22s^22p^1$ – 3 valence electrons
   - C: $1s^22s^22p^2$ – 4 valence electrons
   - N: $1s^22s^22p^3$ – 5 valence electrons
   - O: $1s^22s^22p^4$ – 6 valence electrons
   - F: $1s^22s^22p^5$ – 7 valence electrons
   - Ne: $1s^22s^22p^6$ – 8 valence electrons

2. Lewis dot symbols are a short hand method of representing the valence electrons of a particular atom of an element. Write the Lewis dot symbols for the following atoms: Li, Be, B, C, N, O, F and Ne.

   - Li
   - Be
   - B
   - C
   - N
   - O
   - F
   - Ne

3. There is a relationship between group number from the periodic table and number of valence electrons. What is the relationship?

   Yes. As long as we look at the groups labeled with an “A” after the number (so not the lanthanides, actinides or transition metals), the group number matches the number of valence electrons.

4. Write the Lewis dot symbols for the following elements: Be, Mg, Ca, Sr, Ba, Ra. State the relationship evident here.

   Each of these elements has two valence electrons. They are all in group 2, the alkaline earth metals.
5. Based on position on the periodic table state the number of valence electrons in Se.

Selenium is in the group 6A along with oxygen and sulfur. It has six valence electrons.

Ionic Compounds

Activity 1

The purpose of this activity is to check your understanding of the formation of ions.

1. Metals tend to lose electrons to form cations while nonmetals tend to gain electrons to form anions. Considering that the noble gas electron configuration is the most stable electron configuration, please speculate as to why a metal would tend to lose electrons to become more stable while a nonmetal will tend to gain electrons to become more stable.

Metals only have a few electrons beyond a noble gas configuration. It is more energetically favorable and feasible to lose one or two electrons than to gain six or seven to complete their octets. Nonmetals on the other hand are close to completing an octet and only need to gain a few more electrons to complete their octets.

2. Write the electron configuration for Li atom and Li cation.
   
   Li: $1s^22s^1$
   
   Li$^+$: $1s^2$

3. Draw the Lewis dot structure of Li atom and Li cation.
   
   *Li
   
   Li$^+$

4. Write the electron configuration for F atom and F ion.
   
   F: $1s^22s^22p^5$
   
   F$^-$: $1s^22s^22p^6$
5. Draw the Lewis dot structure of F atom and F ion.

\[
\begin{align*}
\text{F} & : \\
\vdots & \\
\text{F} & : \\
\end{align*}
\]

6. Based on the ions that each of above elements tends to form, what is the predicted ionic compound between Li and F.

\[\text{LiF}\]

Each ion has a charge magnitude of “1.” So for every lithium cation, there would be one fluorine anion.

7. Predict the formula for the ionic compound in formed between magnesium and chlorine.

\[\text{MgCl}_2\]

The chlorine anion has a charge magnitude of “1” but the magnesium cation has a charge magnitude of “2.” So for every magnesium cation, there would be two chlorine anions.

**Activity 2**

The purpose of this activity is to check your ability to form ionic compounds given a metal and a nonmetal and/or a polyatomic ion.

1. Predict the compound formed from rubidium and oxygen. Write the formula and correctly name the compound.

\[\text{Rb}_2\text{O}\]
rubidium oxide

2. Predict the compounds formed from iron and oxygen. Write the formula and correctly name the compound. Notice iron is a transition metal. Write and name the compound for a +2 and a +3 oxidation state of iron. Comment on why there are choices for transition metals and how that is different for main group elements

\[\text{FeO}\]
iron (II) oxide

\[\text{Fe}_2\text{O}_3\]
iron (III) oxide

Many transition metals have multiple oxidation states meaning they have multiple stable cation forms (such as Fe\(^{2+}\) and Fe\(^{3+}\)). Main group elements generally have one stable, possible charge value when an ion forms (such as Ca\(^{2+}\) and Cl\(^{-}\)).
3. Predict the compound formed between the ammonium ion and chlorine. Write the formula and correctly name the compound.

\[
\text{NH}_4\text{Cl}
\]

ammonium chloride

4. Predict the compound formed between calcium and sulfite ion. Write the formula and correctly name the compound.

\[
\text{CaSO}_3
\]

calcium sulfite

**Coulomb’s Law and Lattice Energy**

Lattice energy is defined as the amount of energy necessary to dissociate 1 mole of an ionic crystalline solid into its gas phase ions. Complete this activity to develop the concept that different ionic compounds have different lattice energy.

**Activity 1**

In this activity, we will deconstruct the process of the transformation from pure elements to an ionic crystal to deepen our understanding of ionic compounds and different lattice energies associated with ionic compounds.

1. To better understand the concept of lattice energy think about the formation of an ionic compound from its elements in a step-by-step method by doing a thought experiment. Assume you are making lithium chloride. Write the balanced chemical reaction that describes this change.

\[
2\text{Li} (s) + \text{Cl}_2 (g) \rightarrow 2\text{LiCl} (s)
\]

2. Look up the ionization energy of lithium and write the ionization reaction here. State if this is a step that requires energy or if energy is given off.

\[
\text{Li} (g) \rightarrow \text{Li}^+ (g) + 1\text{e}^-
\]

Ionization energy is about 5.4 eV/atom (520 kJ/mol). This energy is required in order to remove the electron. So energy must be put into the sample to remove the electron.

3. Look up the electron affinity of chlorine. State if this is a step that requires energy or if energy is given off.

The electron affinity of chlorine is \(-3.6\) eV/atom (\(-349\) kJ/mol). Energy is released with the addition of an electron to neutral chlorine.
4. Notice – were these energies computed for solids or gases? If for solid, look up the energy associated with the metal going from the solid to the liquid to the gas phase. Also comment if that process requires energy or if energy is given off.

Values for chlorine should almost always be calculated in the gas phase because that is the most stable phase at room temperature and pressure. The most stable phase for lithium at room temperature and pressure is a solid. So verify that your researched value for the ionization energy for lithium was calculated for gaseous phase lithium.

solid lithium → liquid lithium : 3.00 kJ/mol – energy required
liquid lithium → gaseous lithium : 136 kJ/mol – energy required

5. Now think about a gaseous cation and a gaseous anion. When they come together in space, is energy released or absorbed? In other words is it better energetically for the opposite charges to exist side by side or a long distance apart?

It is energetically more favorable to have the two opposite charges exist side by side. Energy would be released. Right now, two make the two ions, I had to “spend” 5.4 eV for lithium atom and only “earn” 3.6 eV for the chlorine atom.

However, as we bring these charges together, they interact favorably with one another according to Coulomb’s law. The proportionality that describes Coulomb’s law is as follows

\[ E \propto \frac{q_1 q_2}{r} \]

where \( E \) is energy in joules, \( q_1 \) and \( q_2 \) are the charges of the ions in C, and \( r \) is the distance in m.

When the ions are infinitely far apart (\( r \) is very large), no interaction between the two particles (attraction or repulsion) is present. However, as they are brought closer together, more and more energy is released.

6. Now think about a mole’s worth of cations and anions coming together close in space in a crystal lattice, is that more or less energetically favorable than the pairs of gaseous cation and anion?

Now, we are taking the energetically favorable process and expanding it into a 3D structure. So each lithium cation is surrounded by chlorine anions and vice versa. This is very energetically favorable. The positive charges are surrounded and stabilized by negative charges.

7. Try to model the change in energy starting with the free elements going to the crystalline solid.
Activity 2

1. Define lattice energy.

   Lattice energy is the amount of energy required to fully ionize an ionic crystal into its ions in the gas phase.

2. Order the following ionic compounds from lowest lattice energy to highest lattice energy: CaO, CaCl₂, CaS.

   Lowest: CaCl₂ has a lattice energy of about 2170 kJ/mol
   Middle: CaS has a lattice energy of 3006 kJ/mol
   Highest: CaO has a lattice energy of about 3477 kJ/mol

3. Using Coulomb’s law and the words “charge” and “distance”, explain the ordering of the compounds in question 2 above. In your explanation include the importance of ionic radius and use the trend in ionic size from the periodic table to support your answer.

   Again, here is the proportionality that describes Coulomb’s Law:
where $E$ is energy in joules, $q_1$ and $q_2$ are the charges of the ions in C, and $r$ is the distance in m.

Coulomb’s Law shows that if the magnitude of the charges increases so does the lattice energy ($E$ is directly proportional to the product of the charges).

Coulomb’s Law also shows that if the distance between the charges increases, the lattice energy decreases ($E$ is inversely proportional to the distance between the charges). So, differences in ionic radii matter. A smaller ionic radius corresponds with a greater lattice energy. Along the same period, anions are larger than cations because anions have added to their original valence shell but cations have moved down to the next lowest energy level. Within the cations along the same period, the radius decreases from left to right. Within the anions we see the same trend. The ionic radius always increases as you move down a group.

So, let’s consider the examples given in problem #2. The compound with the greatest lattice energy is calcium oxide. The calcium cation has a charge of +2 and the oxide anion has a charge of −2. Big deal – the sulfide anion also has a charge of –2 as well. However, the sulfide anion is larger than the oxide anion (ionic radius increases as you move down a group). Therefore, the distance between the two charges is greater for CaS than for CaO and CaS therefore has a lower lattice energy.

The chlorine anion has a charge of −1. It is on the same period as the sulfide anion but further to the right. This means that the chloride anion should be slightly smaller than the sulfide anion. So why is the lattice energy of CaCl$_2$ lower than that of CaS? The charge of the chlorine anion is half the charge of the sulfide anion. Halving the magnitude of the charge had a greater effect on the lattice energy than the small decrease in ionic radius (distance between charges).

4. **Speculate on the relationship between lattice energy and melting point.**

Lattice energy is the amount of energy required to fully ionize an ionic crystal into its ions in the gas phase. The tighter the ions are bound (due to charge density or distance) the larger the lattice energy.

Melting a substance requires an energy input to break apart the solid structure until the substance is in the liquid phase. While the energy it requires to melt a substance should be lower than the lattice energy (ionizing a substance into the gas phase from its solid crystal), there is a connection. The tighter ions are bound, the higher the melting point. A sample that has a high lattice energy contains tightly bound ions that have strong attractive forces. In order to motivate these tightly bound ions to move and slide past one another more loosely, such a sample must be given sufficient energy by way of heating.
Naming Ionic Compounds

Activity 1

Use this activity to test your knowledge of recognizing formulas and naming ionic compounds. Follow up with creating your own flashcards to make sure you have memorized all the required polyatomic ions. This short activity is just a sample of the types of questions that could be asked on the quiz or the exam.

1. Co(II)Cl₂
   - cobalt (II) chloride
2. Cs₂O
   - cesium oxide
3. Sr(ClO)₂
   - strontium hypochlorite

Activity 2

Use this activity to test your knowledge of writing formulas for ionic compounds. Follow up with creating your own flashcards to make sure you have memorized all the required polyatomic ions. This short activity is just a sample of the types of questions that could be asked on the quiz or the exam.

1. Nickel (I) bromide – NiBr
2. Sodium sulfide – Na₂S
3. Barium carbonate – BaCO₃
4. Potassium phosphate – K₃PO₄