

# Coulomb's Law

## In High School Chemistry

If you teach your students how to apply Coulomb's Law they will be able to:

- Predict and explain Periodic Trends such as atomic radius, ionic radius, and ionization energy.
- Predict trends in lattice energy.
- Use trends in lattice energy to predict solubility of ionic compounds in water.

So, let's start with some simple questions which we will later use Coulomb's Law to answer. Get out a Periodic Table. **Nothing else** – no text, no tables, no calculator.

1. Which atom is larger, Na or K? Explain why.
2. Which atom is larger, Na or Al? Explain why.
3. Which is larger, Na or Na<sup>+</sup>? Explain why.
4. Which is larger, Br or Br<sup>-</sup>? Explain why.
5. Which has a higher first ionization energy, Li or Na? Explain why.
6. Which has a higher first ionization energy, Na or Al? Explain why.
7. Which has a stronger ionic bond, LiCl or NaCl? Explain why.
8. Which has a stronger ionic bond, NaCl or AlCl<sub>3</sub>? Explain why.
9. Which is more soluble (will produce a more concentrated saturated solution) in 80° C water, LiCl or NaCl? Explain why.
10. Which is more soluble in 80° C water, NaCl or AlCl<sub>3</sub>? Explain why.
11. Why is Na<sub>2</sub>O soluble in water while Al<sub>2</sub>O<sub>3</sub> is not?

Simple stuff, right? If you can answer these, using only a Periodic Table, as quickly as you read them then you already know and use Coulomb's Law. If not, here's how it works:

$$\text{Coulomb's Law: } F = k \frac{Q_1 Q_2}{r^2}$$

Hopefully you recognize this as the equation for electrostatic attraction/repulsion, often referred to as the 'inverse square' law. It applies to charged particles, magnets, gravitation, etc. But we are not interested in Forces, we are interested in Energies (such as ionization energy, lattice energy, etc.) So we perform a little calculus and we get:

$$E = k \frac{Q_1 Q_2}{r} \text{ (and since we are not calculating, just comparing, let's simplify to } E \propto \frac{Q_1 Q_2}{r} \text{ )}$$

In English:

$E$  = energy of attraction or repulsion between charged particles

$Q_1$  = charge of first particle

$Q_2$  = charge of second particle

$r$  = distance between charged particles

So: The greater the charges the greater the attraction (if opposite charges) or repulsion (if like charges.)  
And the closer the particles the greater the attraction or repulsion.

In chemistry, we apply Coulomb's Law in 2 situations:

- A. Individual Particles such as atoms or ions. (Ionization Energy)
- B. Ionic Compounds. (Lattice Energy)

The variables take different values in these situations, so let's look at them separately.

#### A. Applying Coulomb's Law to Atoms and Ions

$E$  = ionization energy, the energy needed to remove the outermost electron. The greater the ionization energy, the stronger the attraction and the more energy needed to remove that electron.

$Q_1$  = charge of an electron, -1.

$Q_2$  = effective nuclear charge of protons in nucleus. Attraction for a valence electron is partially shielded by the inner shell electrons. The simple way to estimate effective nuclear charge is to start with the atomic number (total protons) and subtract the number of inner shell electrons. So Hydrogen has an effective nuclear charge of 1 (1 - 0 = 1.) Helium is 2. Lithium is 1 (3 - 2 = 1.) Sodium is 1 (11 - 10 = 1.) All transition metals are 2. All halogens are 7, etc.

$r$  = distance between charged particles which can be approximated by the Period - two atoms in the same period have **approximately** the same atomic radius. Na (Period 3) is smaller than K (Period 4) but similar in size to Al (also Period 3.)

#### B. Applying Coulomb's Law to Ionic Compounds

$E$  = lattice energy, the energy needed to separate oppositely charged ions. The higher the lattice energy, the stronger the ionic bond. The stronger the ionic bond the less soluble in water at a given temperature, since the ions must separate/dissociate from one another and attach to water in order to dissolve.

$Q_1$  = charge of positive ion

$Q_2$  = charge of negative ion

$r$  = bond length, the distance between the nuclei of the 2 ions – can also be approximated by the Period. (Remember to use the Period of the element that is isoelectronic with the ion! Na is in Period 3 and has 3 occupied energy levels.  $\text{Na}^+$  has lost an electron. It has 10 electrons and is isoelectronic with Ne in Period 2.  $\text{Na}^+$  has only 2 occupied energy levels.)

That's all. If you have a Periodic Table and understand how to determine the energy level of the valence electrons then you can answer all of the questions.

**1. Which is larger, Na or K? Explain why.**

K is larger. The size of an atom is determined by the energy level of its' valence electron(s). K has 4 occupied energy levels while Na has 3. Therefore K is larger.

**2. Which is larger, Na or Al? Explain why.**

Na is larger. Both Na and Al have 3 occupied energy levels, so they should be approximately the same size. Adding protons and electrons does not add to the size, since they fill vacant space in either the nucleus or the valence shell. But the attraction for the valence electron in Al (effective nuclear charge  $13 - 10 = 3$ ) is greater than that of Na (effective nuclear charge  $11 - 10 = 1$ .) So the force of attraction is greater in Al and its' valence electrons are pulled a little closer to the nucleus, giving Al a smaller atomic radius than Na.

**3. Which is larger, Na or  $\text{Na}^+$ ? Explain why.**

Na is larger. Na has 11 electrons, so its' valence electron shell is 3.  $\text{Na}^+$  has lost an electron and is isoelectronic with Ne which has only 2 occupied levels which makes it smaller than Na.

**4. Which is larger, Br or  $\text{Br}^-$ ? Explain why.**

$\text{Br}^-$  is larger. Both Br and  $\text{Br}^-$  have the same values for everything in Coulomb's Law, so they are approximately the same size. But even though the added electron in  $\text{Br}^-$  goes into a vacant space in the valence shell, electrons have like charges and repel one another. This increased repulsion between electrons in the valence shell pushes them slightly apart and makes the  $\text{Br}^-$  larger than Br.

**5. Which has a higher first ionization energy, Na or K? Explain why.**

Na has a higher first ionization energy. Both have the same electron charge and effective nuclear charge. But the valence electrons for K are in the 4<sup>th</sup> energy level, while those of Na are in the 3<sup>rd</sup>. So the electrons of Na are held by a greater force of attraction since they are closer to the nucleus. Which is why metallic character (ability to easily lose electrons) generally increases as you move down a family. (Of if you prefer why first ionization energy decreases as you move down a family – its' the same thing.)

**6. Which has a higher first ionization energy, Na or Al? Explain why.**

Al has a higher first ionization energy. Both have the same electron charge, and the valence electrons of both are in the 3<sup>rd</sup> energy level. But Na has an effective nuclear charge of 1 ( $11 - 10 = 1$ ) while Al has an effective nuclear charge of 3 ( $13 - 10 = 3$ .) So the valence electrons of Al are held tighter and it requires more energy to remove them. Which is why metallic character

increases as you move to the left across a Period and decreases as you move to the right. (Or why ionization energy increases as you move to the right across a Period.)

**7. Which has a stronger ionic bond, NaCl or KCl? Explain why.**

The lattice energy of NaCl is stronger, so this is the stronger ionic bond. Both have an electron charge of -1 and an effective nuclear charge of 1. But the valence electrons of Na are in the 3<sup>rd</sup> energy level, leading to a shorter bond length (measured as distance between ionic nuclei) and a stronger ionic bond than KCl.

**8. Which has a stronger ionic bond, NaCl or AlCl<sub>3</sub>? Explain why.**

AlCl<sub>3</sub> has a stronger ionic bond. Both have the same -1 charge for the chloride ion, and both Na<sup>+</sup> and Al<sup>+3</sup> are isoelectronic with Ne and therefore have 2 occupied energy levels. But the higher positive charge of the Al<sup>+3</sup> ion leads to a stronger lattice energy and a stronger ionic bond.

**9. Which is more soluble in 80° C water, NaCl or KCl? Explain why.**

You can dissolve more grams of KCl in 100 grams of 80° C water since it has a weaker ionic bond/lattice energy which requires less energy to separate/dissociate the ions from one another and allow them to attach to the polar water molecules.

**10. Which is more soluble in 80° C water, NaCl or AlCl<sub>3</sub>? Explain why.**

NaCl is more soluble in 80° C water since it has a weaker ionic bond/lattice energy which requires less energy to separate/dissociate the ions from one another and allow them to attach to the polar water molecules.

**11. Why is Na<sub>2</sub>O considered soluble in water while Al<sub>2</sub>O<sub>3</sub> is not?**

Q<sub>2</sub> of oxygen is the same for both compounds, -2. Radius of oxygen is the same for both, and Na<sup>+</sup> and Al<sup>+3</sup> are isoelectronic. So the larger Q<sub>1</sub> charge makes the lattice energy of the Al<sub>2</sub>O<sub>3</sub> greater, and since the ions stay bonded to one another it will not dissociate and dissolve.

FAQs:

Q- Can't I just teach trends, such as atomic radius increases as you move down a family or to the left across a group?

A- Trends are great for quick answers, but I have seen too many students claim that an atom of K is larger than an atom of Na BECAUSE it is lower on the Periodic Table and that shows a total lack of understanding of CHEMISTRY. And how do you explain that as you move to the right across a Period atomic radius gets smaller, even though you are adding protons and electrons, without some reference to Coulomb's Law? Honestly, I would emphasize Coulomb's Law **much** more in Pre-AP courses, since the requirement to explain trends is more of an AP thing. But I do think it is useful in ALL Chemistry courses. 'Atoms are larger to the left within a Period' is **NOT** Chemistry. Coulomb's Law is.

Q- Can I use Coulomb's Law to explain melting/boiling point trends?

A- Probably not. For atoms you have polarizability linked to number of valence electrons able to make a temporary dipole. For ionic compounds you have crystalline structure to deal with which effects the bond length as well. We are also simply estimating effective nuclear charge and ignoring repulsion between positive nuclei when ions get very close or the slight difference between sublevels. Even solubility is greatly effected by temperature. As with (almost) everything in chemistry, we are taking a very simple approach for the first pass and (intentionally) omitting many exceptions.

Q- Is the sign of  $E$  analogous to endo-/exothermic values?

A- I think so. A negative value would require opposite charges. These would produce energy when forming a bond (negative value = exothermic) and require energy to break the bond/separate the oppositely charged particles. Like charges would produce a positive value, indicating an endothermic situation requiring energy to move these particles closer together.

Q- How about comparing KCl to  $AlCl_3$ ?

A- Don't go there. One variable at a time is enough – even in AP Chem students are only asked to compare apples to apples. Leave this one to upperclass/grad Chemistry courses.

Q- What if I have a question? What if my students ask something and I am not sure? I understand, but this is new to me.

A- Email me and I will help.

Shaen McKnight

Richland High School

[shaen.mcknight@birdvilleschools.net](mailto:shaen.mcknight@birdvilleschools.net)