

### Comparing Elements

Elements of same period – Assume radius is same, compare  $Z_{\text{eff}}$

Elements of same group – Assume  $Z_{\text{eff}}$  is same, compare radius

K – Potassium                      Ca – Calcium

Group IA

Period 4

**\*since period is same, assume K and Ca have same radius and compare  $Z_{\text{eff}}$**

$$Z_{\text{eff}} = 19 - 18 = +1$$

Group IIA

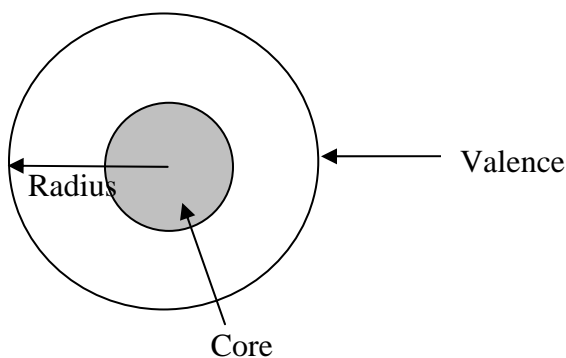
Period 4

$$Z_{\text{eff}} = 20 - 18 = +2$$

$Z_{\text{eff}} = +2$  means that the most loosely bound electron, looking inward sees a +2 charge on core

$Z_{\text{eff}}$  - Effective nuclear charge felt by an electron on the valence shell

Core charge = Effective nuclear charge



**Radius** – How far the most loosely bound electron is from the center (where the nucleus is, at the middle of the core)

**Core has two parts:**

1. Nucleus
2. All shells underneath valence (noble gas configuration) - period 4 and above – these are n-values not periods

Core has two parts, but is viewed as a point charge of magnitude  $Z_{\text{eff}}$  by the most loosely bound electron on the valence.

### **Assumptions**

1.  $S = \#$  electrons in core
2. Electrons in valence shell do not feel repelled by other electrons in valence shell

These are big assumptions and they break down. That is why the  $Z_{\text{eff}}$  model using  $S$  equal to  $\#$  of electrons in core only works to explain general trends, and not the variations that occur within a general trend.

\*Bond – when valence shells share electrons

## **Electron Affinity**

- \*As we look from left to right across the period, the electron affinity of the elements generally becomes more exothermic
- \*Electron Affinity = kJ/mol
- \*Noble gases have positive electron affinity because we cannot easily add an electron, it requires an input of energy (endothermic)
- \*Going down the Group 7A, Fluorine seems out of place
  1. In an ideal trend, F would have an electron affinity of about -375 kJ/mol
  2. The electron affinity for F is less exothermic than -375 kJ/mol because the electrons in F's valence shell are packed closer together than those of Cl, so there is more repulsion when you try to add another electron to F's valence shell

## **The Big Picture: The Pattern of a Coulomb's Law Argument**

- \*Use this slide as a reference for problem solving

## **Map of Chapter 8**

Ionic - Usually a metal and a nonmetal

Molecular – Usually two non-metals

- \*There is some gray area between molecular and ionic bonding

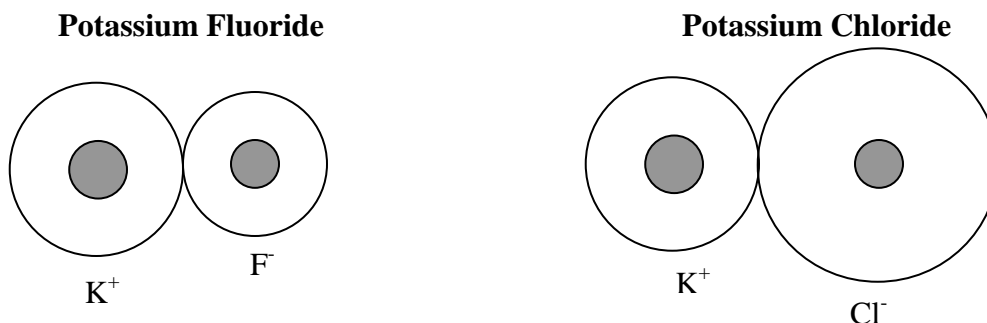
## **Ionic vs. Molecular (what you already know)**

- \*Molecular compounds – Fairly low melting points
- \*Ionic compounds – Exist primarily as solids in nature because bonds are so strong (lattice structure)
- \*Ionic formulas are presented in the simplest repeating unit (NaCl)
- \*Molten ionic compounds – Ions in liquid state (no longer in lattice structure)
  - Needs extremely high temperature to break the ionic bonds holding solid together
- \*Molecular compounds melt at lower temperatures than ionic compounds

## **Lattice Energy in Ionic Compounds is Formation Energy**

\*Lattice Energy – Amount of energy in kJ that it takes to break ionic bonds in one mole of an ionic compound

- \*As the atomic number (radius) of the cation increases, lattice energy decreases
- \*If forming NaCl is exothermic, then breaking it down must be endothermic by the exact same amount
- \*The smallest radius has the smallest distance between two ions
  - Smaller radius = stronger ionic bond = more energy needed to break bond
- \*More shells = less lattice energy



**Which ionic compound has a stronger bond?**

- \*Potassium and Fluoride ions are closer together than Potassium and Chloride ions because the radius of Fluoride is smaller than the radius of Chloride
- \*Potassium Fluoride will have the stronger attraction (stronger ionic bond) because the radius of Fluoride is smaller than the radius of Chloride.
- \*Chloride has more shells, therefore the valence electrons are farther from the nucleus
- \*The ionic bonds in Potassium Chloride are easier to break, so the lattice energy for Potassium Chloride is less endothermic (smaller value)

**Qualitative trends (comparisons) in Strength of Ionic Bonding**

- \*Ionic bonds are stronger than covalent bonds
- \*When bonds take on ionic characters – it is harder to break the bond (polarizing bond)
- \*Remember the definition of “covalent” and rules of bonding will be easy

**Covalent – shared valence electrons**

“co” = shared

“valent” = valence electrons

**Lewis Dot Structure Model**

1. Electrons like to be in pairs
  - a. Shared between two atoms (shared pairs)
  - b. Solely on one atom (lone pairs or unshared pairs)
2. Electron pairs often form in octets around atoms (stable) - known as the octet rule

**Lewis Structures of Simplest Molecules**

- \*From this point we will only worry about valence electrons
- \*Valence electrons are the outermost electrons, so these are the electrons that will interact with other atoms and form bonds
- \*Line = represents a bond (shared PAIR) of electrons
- \*To check that your Lewis structure is correct
  1. Count each electron
  2. Count one bond as one pair (double bond as 2 pair, etc)
  3. Number of electrons should be conserved
- \*Lewis structure always have pairs to form valence shells of eight electrons
- \*Maximum number of bonds = triple bond