

## Chem 115 Notes

04/17/08

- \*Hardest part of Chapter 7 is the wording
- \*Iso = same
- \*Isoelectric = two different things (ion and atom) with same number of electrons
- \*Trend = something that changes when something else changes (fashion trend = skirt lengths change as time changes)
- \*A trend is stated, for example: As \_\_\_\_\_ increases, \_\_\_\_\_ decreases.
- \*Increase = going up
- \*Decrease = going down

### Summary of Ionization Energy Trends

- \*Ionization Energy increases as you move from left to right across the periodic table
- \*Ionization Energy decreases as you move from top to bottom of a group in periodic table

### Ionization Energies of Elements (noble gases)

- \*Ionization Energy – Energy required to remove the most weakly bound electron from an atom or ion
- \*2<sup>nd</sup> Ionization Energy – After 1<sup>st</sup> electron is removed how much energy is needed to remove next electron
- \*Trend – As atomic number increases, the 1<sup>st</sup> ionization energy of noble gases decreases

### Comparing Noble Gas Group Elements

- \*Noble gases – valence shells are full ( $s^2p^6$ )
- \*What is the Same - valence shells are full ( $s^2p^6$ )
- \*What is Different – n values
- \*How does the Difference explain the trend seen?
  - As atomic number increases, the first ionization energy of the noble gases decreases because the electron is farther away from the nucleus.
- \*Less energy is needed to pull an electron off if it is further from the nucleus

### Ionization Energies of Elements (alkali metals)

- \*Trend seen: As atomic number increases, the first ionization energy of the alkali metals decreases.

### Comparing Alkali Metal Group Elements

- \*How are the Alkali Metals the same?
  - All end in  $s^1$
- \*How are the Alkali Metals different?
  - n value of  $s^1$  is different ( $1 s^1, 2 s^1, 3 s^1$ )
- \*K  $\rightarrow$   $K^+ + e^-$
- \*Rb  $\rightarrow$   $Rb^+ + e^-$

## Ionization Energies of Elements (Alkali Metal Ions)

\*When you pull off an electron the Alkali Metals have a complete shell

## Comparing Alkali Metal Group Ions

\*How are they the same?

- Full valence shells ( $s^2p^6$ )

\*How are they different?

- n values

\*How does this difference explain the trend seen?

- As atomic number increases, the second ionization energy of the alkali metals decreases because there are more shells, so electrons are farther from the nucleus.

## What We've Observed So Far

\*Atomic number increases as 2<sup>nd</sup> ionization energy decreases

\*Number of shells are different → energy is different

\*Trend on group → use argument that shells are different

## Core vs. Valence

\* $Z_{\text{eff}}$  – Electron on valence shell looking inward, what attractive charge do you see that holds you onto the atom (or ion)?

- Nucleus (+ charge) is shielded by core electrons (fewer - charges than the nuclear charge)

\*  $Z_{\text{eff}} = Z$  value – screening number

\*Z value = atomic number

\*Screening number = noble gas core

## Example: Magnesium Atom

Z value = 12

Screening number = 10 (which is atomic number of Ne, the noble gas configuration underneath)

$Z_{\text{eff}} = 12 - 10$

$Z_{\text{eff}} = 2$

\* $Z_{\text{eff}}$  - also shown in the Roman numerals at top of periodic table

## Recall Coulomb's Law

\*Ionization Energy of Neon : Ionization Energy of Sodium

\*Na = larger distance from nucleus = force smaller

\*Easier to remove and  $e^-$  from Na than Ne

Differences to look for when making an argument

1. Shells

2.  $Z_{\text{eff}}$

\*As  $Z_{\text{eff}}$  increases, the force holding electron increases

\*As shells increase, the force holding electron decreases

## Example of Building a Chemical Explanation

### Neon (Ne)

Core =  $1s^2$

Valence =  $2s^2 2p^6$

### Sodium (Na)

Core =  $1s^2 2s^2 2p^6$

Valence =  $3s^1$

\*Easier to remove an electron from Na

1. The valence electrons of Na are farther from the nucleus
2.  $Z_{\text{eff}}$  of Na = +1  
 $Z_{\text{eff}}$  of Ne = +8

Comparing Na and Ne ionization energies, both shells and  $Z_{\text{eff}}$  work in same way to affect Ionization Energy

## Atomic Radius Comparisons

- \* The atomic radius is determined by the distance from the nucleus of the most loosely bound electron
- \* Moving from left to right across the periods, atomic radius decreases as  $Z_{\text{eff}}$  increases (force holding the mostly loosely bound electron is stronger)
- \* Moving from the top to the bottom of a group, atomic radius increases...more shells (force holding the most loosely bound electron is weaker)
- \* Force and radius are inversely related (atomic number increases, atomic radius decreases)

## Radii of Cations and Anions vs. Neutral Atoms

### Cations and Anions

- \* Moving from top to bottom in group, radius increases (shells argument)
- \* Moving from left to right in a period, radius decreases for cations
- \* Moving from left to right in a period, radius increases for anions (because adding electrons which repel other electrons that are already there)

## Electron Affinity

- \* Electron Affinity is the opposite of Ionization Energy
- \* Electron affinity is the energy required to add an electron onto an atom
- \* In most cases it is Exothermic (releases energy), therefore has a (-) value
- \* Values listed on slide are KJ of E
- \* F is more exothermic than Li (larger negative value)

## Electron Affinity

- \* Circled are the exceptions = positive values (need to add energy to add  $e^-$ )
- Why are they exceptions?
- \* Difficult to add  $e^-$  to Helium because it is stable ( $e^-$  will need to go to next shell)
- \* N = has a half filled shell so it has some stability
- \* Ne = stable full shell, doesn't want another electron, so must add energy (endothermic) to force it to accept another electron

## Building a Chemical Explanation

Explain why Br has a smaller (less exothermic) electron affinity than Cl:

- \* Br and Cl are both halogens (same  $Z_{\text{eff}}$ )
- \* Will Cl or Br accept an  $e^-$  more easily?

- \*Cl will because it can get closer to the core (closer to  $Z_{\text{eff}}$ )
- \*Charges in valence are same, so the difference must be in distance
- \*Chlorine = smaller radius = more exothermic

**Answer to homework problem 7.34**

a) Why is  $\text{Cl}^-$  larger than  $\text{Cl}$ ?

1. First identify that the question is asking about radius
2.  $\text{Cl}^-$  has more electrons, therefore it is larger
3. Radius is larger due to e- repulsion



b) Why is  $\text{S}^{2-}$  larger than  $\text{O}^{2-}$ ?

\*Sulfur ion is larger than oxygen ion because sulfur has more shells

1. Have same  $Z_{\text{eff}}$
2. Different # shells
  - \*S has 3 shells
  - \*O has 2 shells

c) Why is  $\text{K}^+$  larger than  $\text{Ca}^{2+}$ ?

1. Same period (4)
2.  $Z_{\text{eff}}$  of  $\text{K}^+ = 19 - 18 = +1$
3.  $Z_{\text{eff}}$  of  $\text{Ca}^{2+} = 20 - 18 = +2$
4. Larger  $Z_{\text{eff}}$  is  $\text{Ca}^{2+}$  = smaller radius on  $\text{Ca}^{2+}$
5.  $\text{K}^+$  is larger