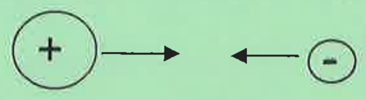
Name Key

Periodic Trends & Coulomb's Law

Coulomb's Law gives us a way to measure the attractive force that comes from the separation of charges. You know that a positive charge and a negative charge attract. Think of a positive charge and a negative charge near each other.



Each charge is attracted to the other and will move towards it. If the charges are moving, they must have a force acting on them. Coulomb came up with a way to measure the attractive force between two charges. His equation is:

$$F = \frac{kq_1q_2}{d^2}$$

- **F** is the force of attraction/repulsion in newtons (N)
- **q_1** and **q_2** are a measure of the charge on each particle. For example, the charge on the electron is -1, so q = -1 for the electron.
- **d** is the distance between the two charges measured in meters.
- **k** is a proportionality constant.

Critical Thinking Questions

1. According to Coulomb's Law, what happens to the force as the charges get farther apart?

   Force (attraction or repulsion) decreases

2. According to Coulomb's Law, what happens to the force as the charges get larger?

   Force (attraction or repulsion) increases

3. If q = -1 for an electron, what is q for a proton? +1

4. What is q for a neutron? 0

5. What is q for the nucleus of a carbon atom, as felt by the 1s electrons? +6 (no shielding of 1s)

6. What is q for the nucleus of a phosphorous atom, as felt by the 2p electrons? +13 (+15 - 2 shielding)

7. If the attractive force between an electron and its nucleus is 10 N and in another atom it is 5 N, which will have a higher ionization energy? 10 N ≠ greater attraction (Technically a −F = attraction) But, assuming comparing -10 N to -5 N, -10 N represents ↑ attraction = ↑ IE

8. If a nucleus has more protons, will the electrons feel a greater or lesser attractive force? Use Coulomb's law to defend your answer. Greater, increasing #P will increase q_1, increasing force

9. If the nucleus and electrons are closer in one atom than another, will the electrons feel a greater or lesser attractive force. Use Coulomb's law to defend your answer. Greater, decreasing distance decreases "d", increasing force

CHEMTIVITY 2: RADIUS, EN, IE AND CHEMICAL PROPERTIES

Use your understanding of the trends for radius, electronegativity, and ionization energy to answer the following:

1. Define ionization energy

   Energy required to remove an electron from an atom

2. As you go across a period on the periodic table, in general, what happens to ionization energy? Explain why.

   Increases due to an increase in effective nuclear charge (smaller radii).

3. What happens to atomic radius across a period on the periodic table? Explain why.

   Decreases, see #2

   +3 2e- 1e-   +4 2e- 2e-   +5 2e- 3e-

   Li   Be   B

   ENC = +1   ENC = +2   ENC = +3

4. Using Coulomb's Law, explain how atomic radius and ionization energy are related.

   Smaller radius = ↑ IE

   small radius = small distance b/w p^+ & e^-, causing force of attraction to increase per Coulomb's Law

   $$F = \frac{k q_1 q_2}{d^2}$$

5. Remember that as you go down a group on the periodic table, atomic radius gets larger. Explain the reasoning for this trend.

   Addition of energy levels (no change in ENC)

   +3 2e- 1e-   ENC = +1

   Li

   +11 2e- 8e- 1e-   ENC = +1

   Na

6. What would you predict for the trend in ionization energy as you go down a group on the periodic table? Explain your reasoning.

   IE decreases down a group since radius increases, causing less attraction b/w p^+ & e^- per $F = \frac{k q_1 q_2}{d^2}$

7. Define electronegativity.

Ability of an atom to attract another atom's valence electrons

8. Predict the trend in electronegativity across a period and down a group. Explain why these trends occur.

→ Period: ↑ EN, radius is decreasing allowing the atom to get its nucleus (p^+) closer to the e^- it is trying to attract

↓ Group: ↓ EN, radius is increasing causing nucleus (p^+) to be farther from e^- it is trying to attract

9. Using your knowledge of the quantum mechanical model, explain why the noble gas configuration is stable.

- Half-full sublevels (ie ↑ ↑ ↑ 2p) gain stability by reducing electrostatic repulsion b/w e^- by spacing into individual orbitals.
- Full sublevels (like nobel gases) gain stability by counteracting electrostatic repulsion w/ magnetic field attraction (all e^- are paired w/ opposite spins & therefore opposing magnetic poles)

10. Explain how your answer in #9 is shown in the electronegativity and ionization energy of noble gases.

Their stability is reflected in their EN & IE values:

- EN = 0 ; don't attract / gain e^-
- IE = huge ; don't lose e^-

Not gaining or losing e^- = not reactive = stable

11. Using your knowledge of how metals and non-metals behave during bonding, which (electronegativity/ionization energy) would be a good predictor of reactivity for each? Explain.

Metals = ↑ radius = ↓ IE = lose e^-

Non-Metals = ↓ radius = ↑ EN = gain e^-

12. Based on the trend in ionization energy, predict the trend in reactivity for the alkali metals. Explain.

Reactivity should increase as you go down the alkali group b/c ↑ radius = ↓ IE = more likely to lose e^- (react)

13. Based on the trend in electronegativity, predict the trend in reactivity for the halogens. Explain.

Reactivity of halogens should decrease down a group b/c

↑ radius = ↓ EN = less likely to gain e^- (less reactive)

14. Select the pair of elements that reacts most readily?
   a. Li + Br_2
   b. Li + Cl_2
   c. K+Br_2
   d. K+Cl_2

Li more reactive than K? (no, smaller radius = bigger IE)

Br more reactive than Cl? (no, larger radius = small EN)

K & Cl most reactive (willing to lose/gain e^-)

15. Without referring to a reactivity series, predict the spontaneity of the following reactions. For those that occur, give the products.

   a. Na + LiCl → NaCl (aq) + Li (s)   Na more reactive (↓IE), able to displace Li

   b. Na + KCl → no rxn

   c. Br_2 + KCl → no rxn

   d. Br_2 + KI → KBr (aq) + I_2 Br more reactive (↑EN) than I, able to displace

AP Formatted Questions

1. Answer the following questions about nitrogen, oxygen, fluorine and iodine using principles of atomic and molecular structure.

Element	First Ionization Energy ($kJ\ mol^{-1}$)
N	1402
O	1314
F	1681
I	???

CL = coulomb's law

$F = \frac{k q_1 q_2}{d^2}$

a) Explain why N has a smaller first ionization energy than F.

F has smaller radius = ↑ force of attraction per CL = ↑IE

b) Explain why O has a smaller first ionization energy than N.

N has stable 2p sublevel due to half-filled configuration, therefore more energy is required to take it out of the stable config.

c) Would you predict the first ionization energy of iodine to be greater than, less than, or equal to fluorine? Explain.

Less, I has much larger radius = ↓IE per CL

d) Which of the atoms listed above would have the largest atomic radius? Explain.

I, N/O/F all have 2 energy levels occupied by e^-
I has 5.

e) When bonded with fluorine, nitrogen atoms form the molecule NF_3; however atoms of iodine can form IF_3 and IF_5 molecules. Explain.

I can form an expanded octet due to the presence of d-sublevel e^-, N has no d block e^- to utilize in bonding (covered more later, but d block e^- allow for formation of trigonal bipyramidal ~~[illegible]~~ geometry)

f) Regarding question (e), why are the formulas predicted to be NF_3 and IF_3 rather than F_3I and F_3N?

More electropositive element is listed first (likely to lose e^-), F is more electronegative than both N & I. (greater than N due to smaller radius caused by ↑ENC
greater than I due to smaller radius caused by fewer energy levels)

2. Answer the following questions using principles of atomic and molecular structure. The elements in the table below (W, X, Y, and Z) are actual elements found in either period 2 or 3 in the periodic table.

Element	First Ionization Energy (kJ mol^{-1})	Second Ionization Energy (kJ mol^{-1})	Third Ionization Energy (kJ mol^{-1})	Fourth Ionization Energy (kJ mol^{-1})
W	520	7298	11815	——
X	900	1757	14850	21000
Y	801	2427	3660	25000
Z	496	4562	6910	9543

a) Which of the elements listed above has a valence electron configuration of $3s^1$? Justify your answer.

Z

Both W & Z exhibit a large increase in IE after removing the first e^-, suggesting both have s^1 valence e^- configs. Since W's first IE is greater than Z's, it is likely $2s^1$ (greater attraction) while Z is $3s^1$

b) Which of the elements listed above is an alkaline earth metal? Explain.

X

Alkaline earth metal = group 2 = s^2 valence configuration

Since X experiences a dramatic increase in IE after losing 2 e^-, it likely is a metal w/ 2 valence e^- that is isoelectric w/ a noble gas after losing 2 e^-

c) Which of the elements listed above has the largest atomic radius? Explain.

Z

- Largest radius = Smallest IE
- Previously determined W is in 2nd period, Z is in 3rd, ∴ W should be smaller radius
- Between Z, Y, & X; Z has smallest IE = largest radius (↑d = ↓F attraction)

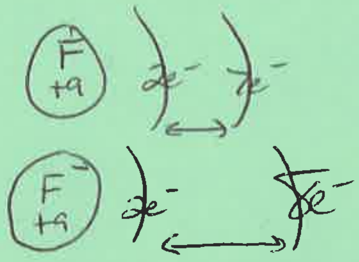
d) Which element, X or Y, has more protons? Assume both have the same principal valence energy level, *n*.

Y

- If both are in the same energy level/period, the atom with the larger nuclear charge will have a higher IE, (shielding & distance due to energy levels are the same) but we don't know for sure, because that trend is not continuous across a period (blips up & down)
- We can conclude that X is in group 2 (large increase in IE after losing 2nd e^-) & Y is in group 3 (large increase in IE after losing 3rd e^-) due to the IE jumps and X having a slightly larger 1st IE than Y (removing e^- from full s sublevel requires more energy than removing single e^- from p sublevel)
- Y therefore likely has more protons due to its periodic placement

3. Use appropriate principles of atomic structure to account for each of the following statements. Be sure to discuss *both* substances in your response.

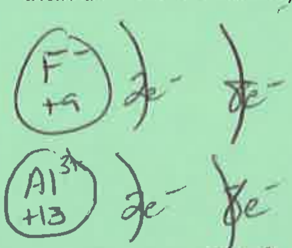
a) The radius of the fluorine ion, F^-, is larger than the atomic fluorine, F.



• Both atoms have identical ENC & energy levels

• The addition of an electron to form F^- creates additional repulsion b/w the 1st & 2nd energy level, increasing the radius

b) The Fluorine ion, F^-, and the aluminum ion, Al^{3+}, are isoelectric. However, the fluorine ion, F^-, is larger than the aluminum ion, Al^{3+}.



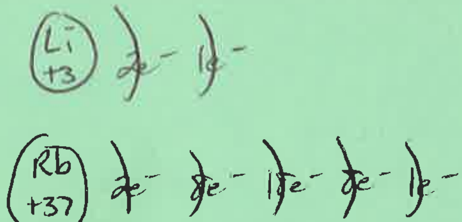
• Both have identical Energy levels & shielding

• The larger nuclear charge of Al^{3+} gives it a larger ENC, decreasing its radius

The table below compares the elements lithium, sodium, potassium and rubidium.

Element	Atomic Radius (pm)	Difference in Radii (pm)
Li	145	35
Na	180	
K	220	15
Rb	235	

c) Rubidium has a larger atomic radius than lithium.



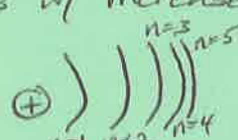
• Both have same ENC (+1)

• Rb has more principal energy levels (5) than Li, physically increasing the distance b/w the nucleus & valence level.

d) The difference between atomic radii of lithium and sodium is relatively large compared to the difference between the atomic radii of potassium and rubidium.

According to Coulomb's Law, the distance b/w particles highly influences the force of attraction b/w the particles as it is squared.

Therefore, energy separation at higher energy levels is smaller than at lower levels due to the attractive force b/w p^+ & e^- dissipates w/ increased distance



$$F = \frac{kq_1q_2}{d^2}$$

4) Refer to the table on the right for the following questions:

(A) O
(B) Ba
(C) Rb
(D) Mg
(E) P

a) Is the most electronegative. A

b) Has the greatest first ionization energy. A

c) Has the largest radius. C — ok, so in real-life Rb > Ba. But if you follow the trends, Ba should be larger w/ more energy levels

d) Is the most reactive metal. C

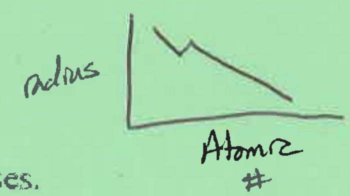
e) Is the most reactive non-metal. A

5) Which species listed are isoelectric? D

(A) Ba, Ra, Ca
(B) Cl^{-}, As^{3-}, S^{2-}
(C) Cr, Mn, Fe
(D) S^{2-}, Cl^{-}, K^{+}
(E) Ba^{2+}, Ca^{2+}, Be^{2+}

6) In the periodic table, as the atomic number increases from 3 to 10, which statement best describes the effect on atomic radius? B

(A) It increases.
(B) It decreases.
(C) It remains constant.
(D) It decreases and then increases.
(E) It increases and then decreases.



7) Which of these sequences refer to a correct trend in ionization energy? ________

I. Al < Si < P < Cl ✓
II. Be < Mg < Ca < Sr ✗
III. I < Br < Cl < F ✓
IV. $Na^{+} < Mg^{2+} < Al^{3+} < Si^{4+}$ ✓

(A) I only
(B) III only
(C) I and II only
(D) I and IV only
(E) I, III, and IV only

8) Based on the information in the table below, what is the most likely charge of element X? C

Element	First Ionization Energy ($kJ\ mol^{-1}$)	Second Ionization Energy ($kJ\ mol^{-1}$)	Third Ionization Energy ($kJ\ mol^{-1}$)	Fourth Ionization Energy ($kJ\ mol^{-1}$)
X	900	1757	14850	21000

Jump = group 2 metal

(A) 1–
~~(B) 1+~~
(C) 2+
~~(D) 3+~~
(E) 4+

9) The effective nuclear charge experienced by the valence shell electron of Na is different than that experienced by the valence shell electron(s) of Ar. This difference best describes which of the following? Na has D

(A) a smaller radius than Ar. ×
(B) a higher melting point than Ar. ×
(C) a greater electron affinity than Ar. ×
(D) a lower first ionization energy than Ar ✓
(E) a higher neutron-to-proton ratio than Ar. ×

Na +11 2e⁻ 8e⁻ 1e⁻ ENC=+1

Ar +18 2e⁻ 8e⁻ 8e⁻ ENC=+8

10) Which group of elements has nearly the same atomic radius? E

(A) F, Cl, ~~I~~
(B) Li, C, ~~F~~
(C) N, S, ~~Br~~
(D) Li, K, ~~Cs~~
(E) Mn, Fe, Co

11) Which of the following elements below is least reactive? C

(A) F
~~(B) Li~~
(C) Ne
~~(D) Xe~~ ← same EN / reactivity
(E) Ra