The chemical elements are fundamental building materials of matter, and all matter can be understood in terms of arrangements of atoms. These atoms retain their identity in chemical reactions.

Big Idea #1 : Atomic Structure

Enduring Understanding

- 1.A: All matter is made of atoms. There are a limited number of types of atoms: these are the elements.
- 1.B: The atoms of each element have unique structures arising from interactions between electrons and nuclei.
- 1.C: Elements display periodicity in their properties when the elements are organized according to increasing atomic number. This periodicity can be explained by the regular variations that occur in the electronic structure of atoms. Periodicity is a useful principle for understanding properties and predicting trends in properties. Its modern-day uses range from examining the composition of materials to generating ideas for designing new materials.
- 1.D: Atoms are so small that they are difficult to study directly; atomic models are constructed to explain experimental data on collections of atoms
- l.E: Atoms are conserved in physical and chemical processes. structured to explain experimental data on collections of atoms.

Nucleus

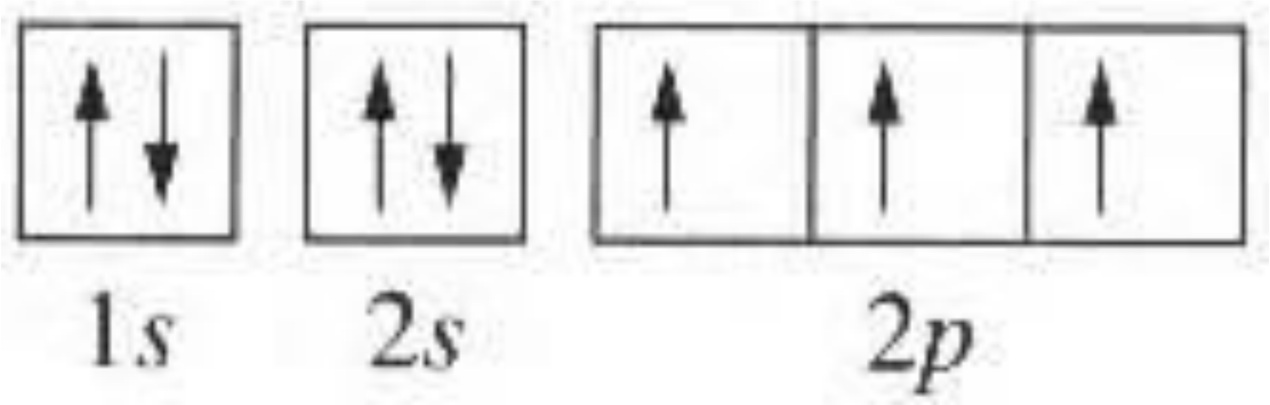
- Number of protons = atomic number = determines the element
- Mass number: total mass of a single atom (p^+ + n^1_0 = total mass (amu))
- Isotopes: Due to varying number neutrons atoms of the same element can vary in mass and therefore weight.
- Average atomic mass:
  - Average mass (amu) of each atom naturally occurring on earth.
  - This number also correlates to 1 mole for number of grams.
- Nucleus is very small and very dense

Electron cloud

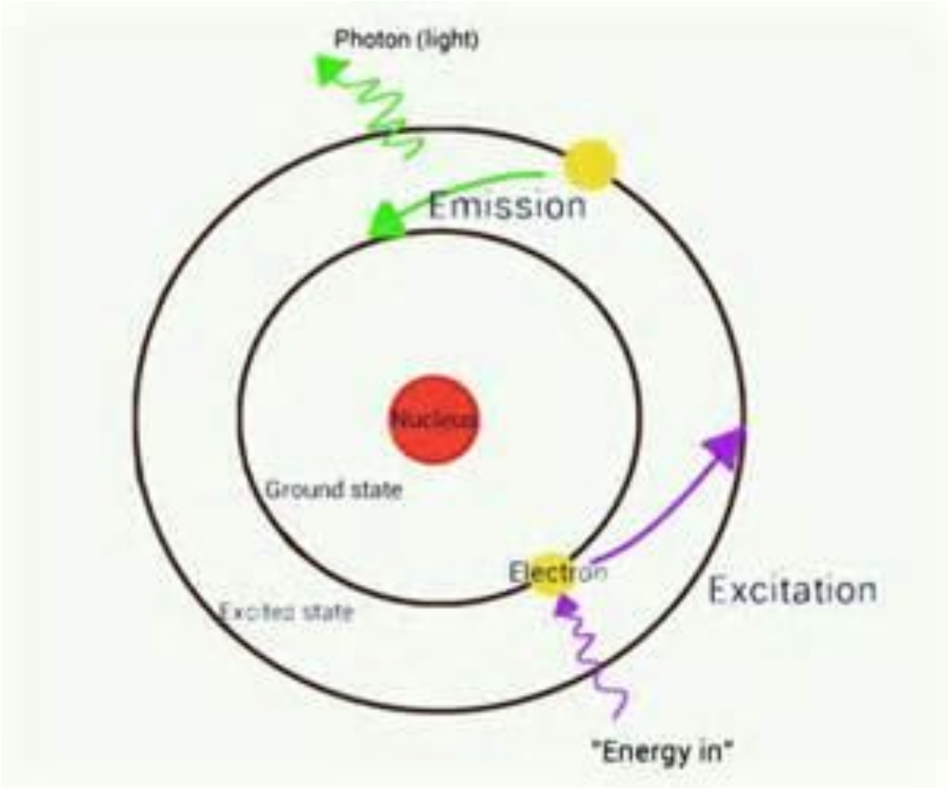
- Electrons control the bonding of atoms, more specifically valence electrons.
- Electron configurations:
  - model to give description of electron location and eventually some of properties.
  - n = energy level n = 1, 2, 3, 4, 5, 6, 7 (principle quantum number)
  - l = orbital type l = s, p, d, f (Orbital quantum number
  - l_m= Sub orbital type s = 1(0) p = 3 (+1, 0, -1) d = 5 (+2, +1, 0, -1, -2) f = 7 (magnetic quantum number)
  - l_s = ↑ spin (+1/2) ↓ spin(-1/2) (spin quantum number)

Hund’s rule

- Electrons are lower energy (more stable) if they are in separate sub orbital rather then paired in a single orbital.



Ground state vs. Excited state

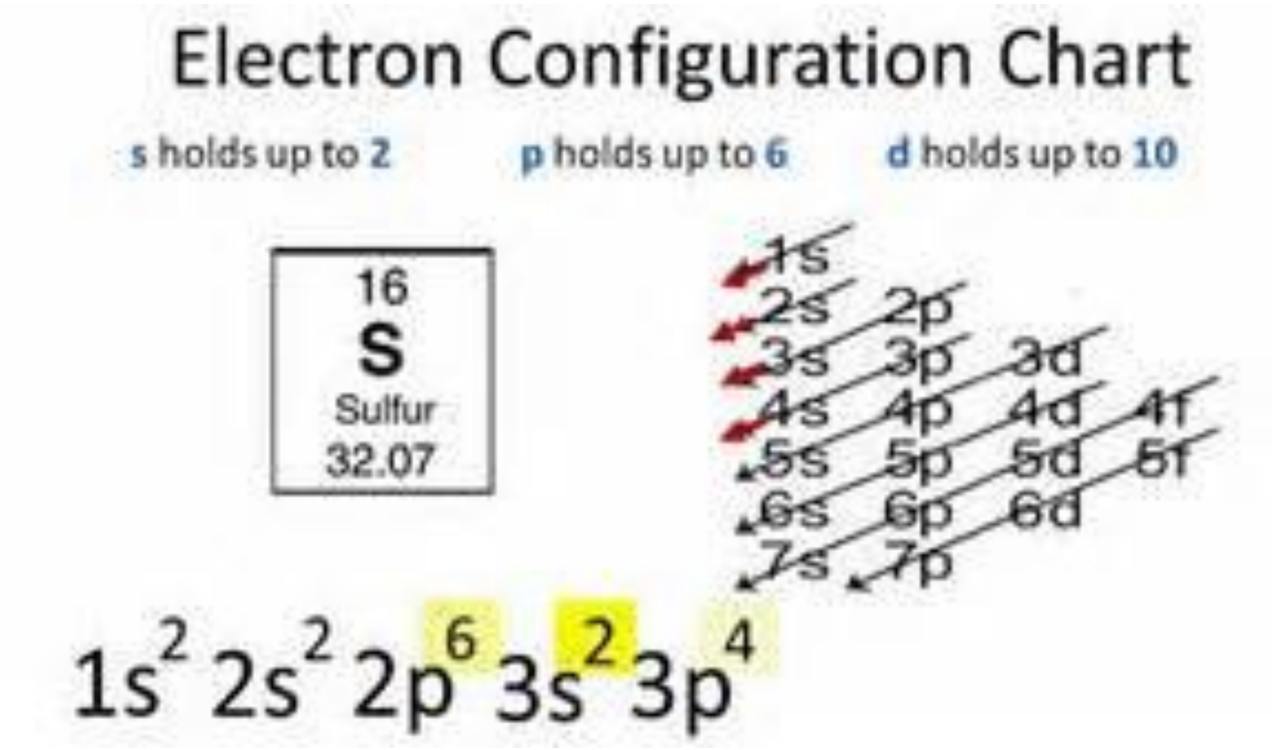


Note: electrons can absorb energy and get pushed up into
Higher energy shells. They will fall back down releasing energy
In the form of a photon.

- This photon is characteristic of the of the structure of the atoms Electrons and therefore the atom itself.

NEARLY EVERY ELECTRON CONFIGURATION IS DONE FOR GROUND STATE ELECTRONS.

Electron configuration example



Pauli Exclusion Principle

- No two atoms can have the same quantum numbers.

Periodic Trends

- This might be the most tested aspect of Big Idea #1

Trends:

- Atomic radius : Size of an atom.
- Ionization energy: Amount of energy needed to be invested in order to remove an electron. Energy is added so this is endothermic.
- Electron affinity: Energy released when an electron is added to an atom. Bringing a + closer to a – is energy releasing or exothermic.
- Electronegativity: atoms attraction for electrons in a bond

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

- Coulombs law is the justification for nearly all of the atomic trends.

- As charged particles get closer together the attraction gets larger
- Charged particle (+) attract (-)
  - + nucleus (charge builds as protons are added)
  - - electrons (are spread out and act as single negative charges)
- Particle (-) repel (-)
  - - electrons are negative

Atomic radius

- What happens when you add just a proton?

Justify with Coulombs law?

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

Example: Which atom is bigger Oxide or Fluoride?

O^{-2} vs. F^{-1}

Notice these atoms have the same # of electrons but F^- has one more proton!

Atomic radius

- What happens when you add just a proton?

Justify with Coulombs law?

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

Adding more charge will cause the atom to become smaller. Greater charge = greater attraction.

$F^{-1} < O^{-2}$ Due to larger number of protons increased.

Atomic radius

- What happens when you add an electron? Use Coulombs law to justify.

Ex)

Oxygen vs. Oxide

O vs. O^{-2}

Ex)

Potassium vs. Potassium ion

K vs. K^{+}

Atomic radius

- What happens when you add an electron? Use Coulombs law to justify.
- Electrons repel each other if they are with in a shell making it larger
- If a new shell gets opened or closed the size will change dramatically!

Ex)

Oxygen vs. Oxide

O vs. O^{-2} 2 electrons are added to the same energy level = slightly larger.

$O < O^{-2}$ due to repelling electrons!

Ex)

Potassium vs. Potassium ion

K vs. K^+ 1 electron is removed eliminating need for entire energy level.

Significantly smaller! $K^+ <<<< K$

Atomic Radius

- What happens if you add a proton and an electron
  - Force increases :

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

- Ex) Sulfur vs. Chlorine and Argon vs. Potassium

Atomic Radius

- What happens if you add a proton and an electron
  - Force increases :
    - Radius gets smaller if added to same energy level do to Force↑
    - Radius will get much bigger if you add an e^- that opens up a new energy level!

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

- Ex) Sulfur vs. Chlorine and Argon vs. Potassium

S > Cl

K >>>> Ar

Overall trend

Ionization energy

- Amount of energy needed to be added to the system to separate a e^- from a + nucleus.
- Always adding energy…. Endothermic
- Justification is always based upon Coulombs law.

- Example: Highest ionization energy?

    K vs. Ar

    K vs. Ca

    O^{-2} vs. O

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

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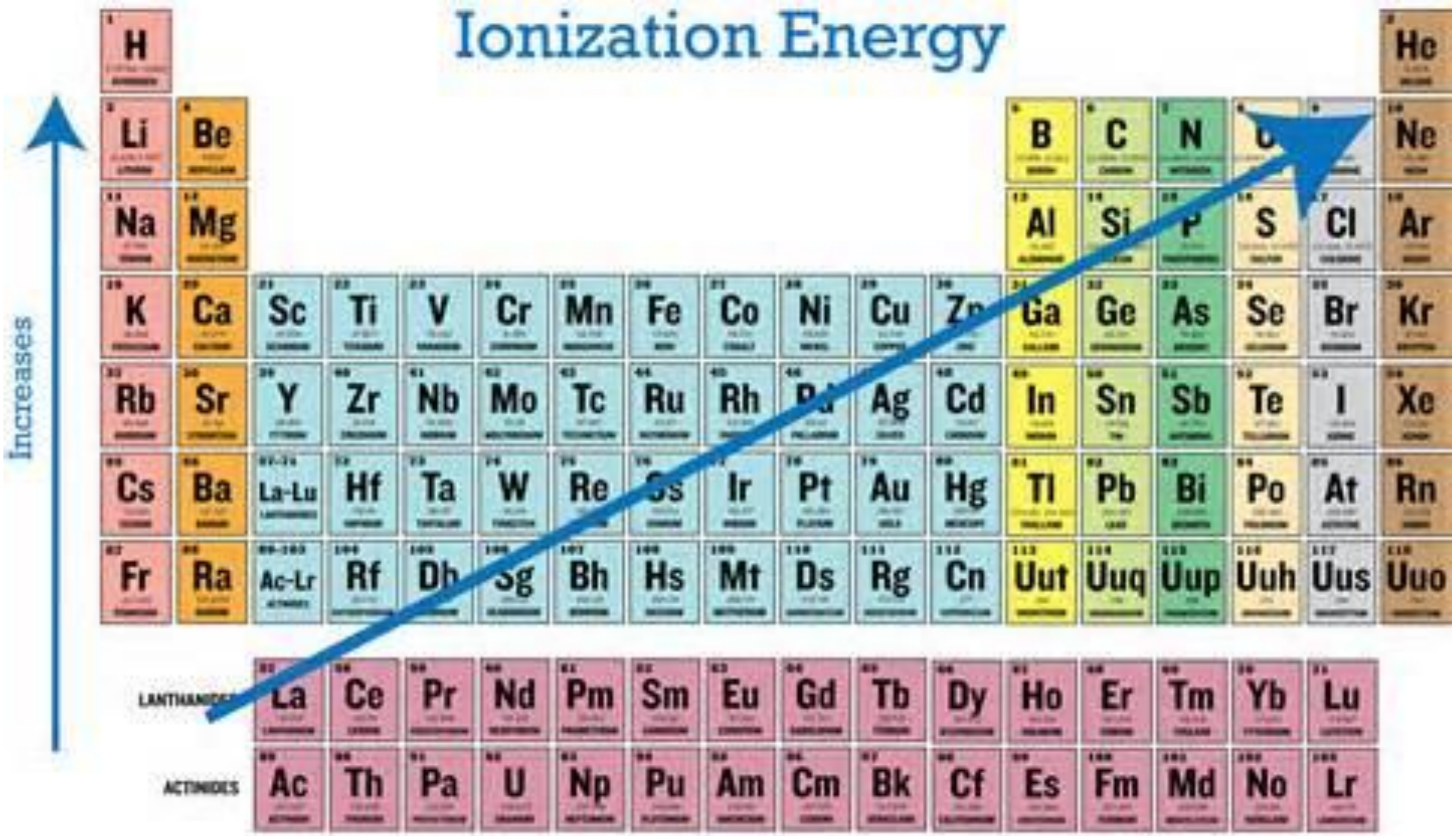
- Example: Highest ionization energy?

K vs. Ar = Ar >>> K due to much smaller size

K vs. Ca = Ca > K due to slightly smaller size

O^{-2} vs. O = O > O^{-2} due to the fact that O^{-2} is slightly bigger

Ionization energy: Correlates inversely to radius



Electron affinity

- How much energy is released when you bring an electron closer to a + nucleus.
- Remember + and – naturally attract, NO energy needs to be invested here in fact energy gets released.
- Exothermic!

Which of the following reactions are Exothermic or endothermic?

$Na \Rightarrow Na^{+} + 1e^{-}$ $\Delta H = ?$ (+/-)

$F + 1e^{-} \Rightarrow F^{-}$ $\Delta H = ?$ (+/-)

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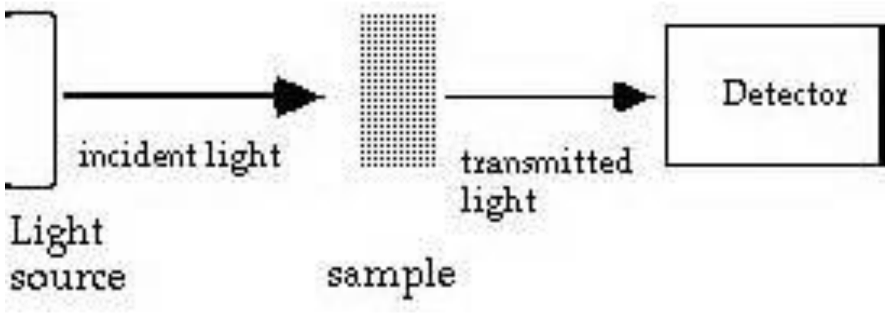
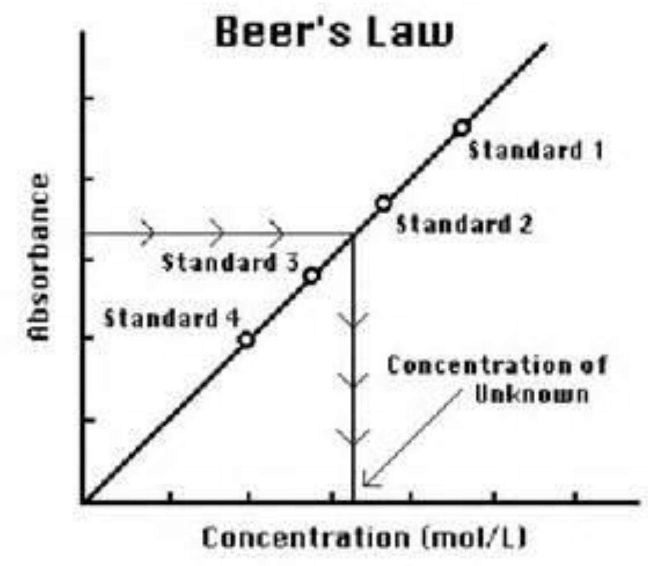
$F + 1e^{-} \Rightarrow F^{-}$ $\Delta H = -$ electron affinity

1.D: Atoms are so small that they are difficult to study directly; atomic models are constructed to explain experimental data on collections of atoms.

- Spectroscopy : Using light as a means to study substances
  - Beer’s law:
    - Concentration of light getting through a solution is related to the concentration.



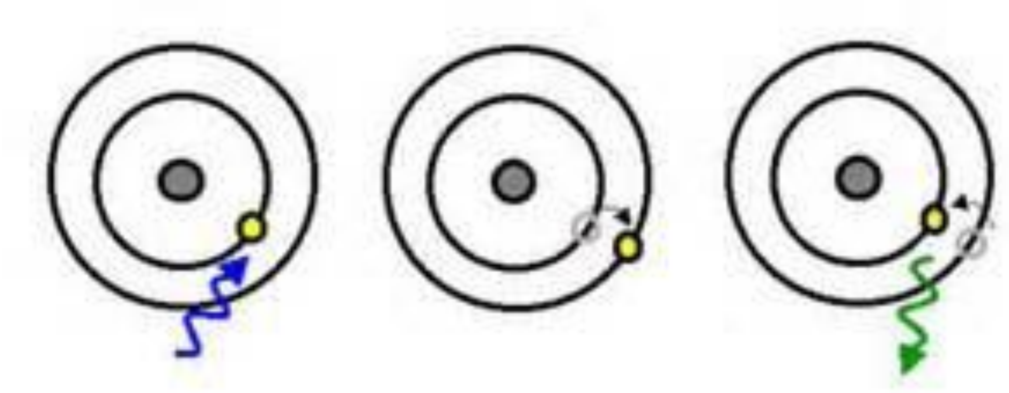


Absorbance = abc (Rarely or never used)
a = absorptivity constant
b = distance across sample
c = concentration

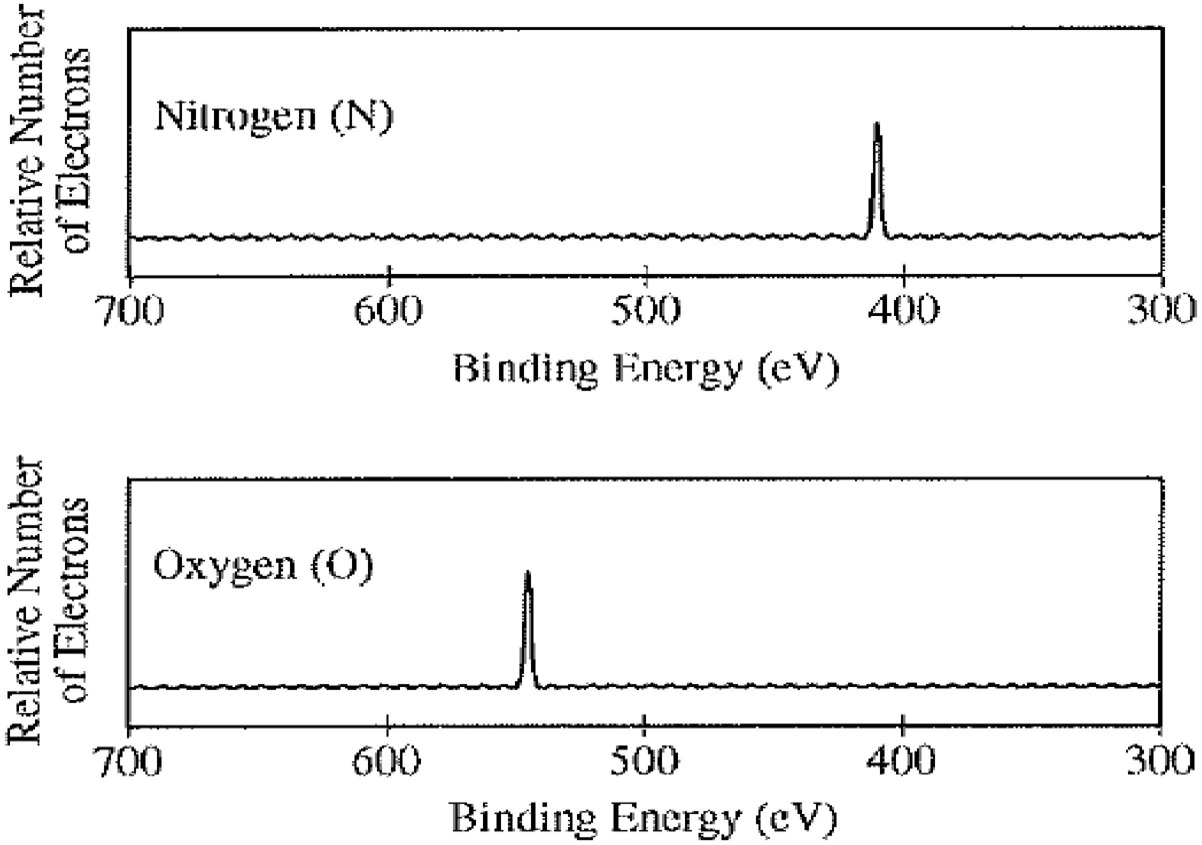
Flame test/ IR spectroscopy

- Type of spectroscopy
  - Add energy to a substance and record the type of light that is being released.
  - Not all light will be visible.

  - Example:
    - IR spectroscopy
    - Flame test



Photoelectron spectroscopy

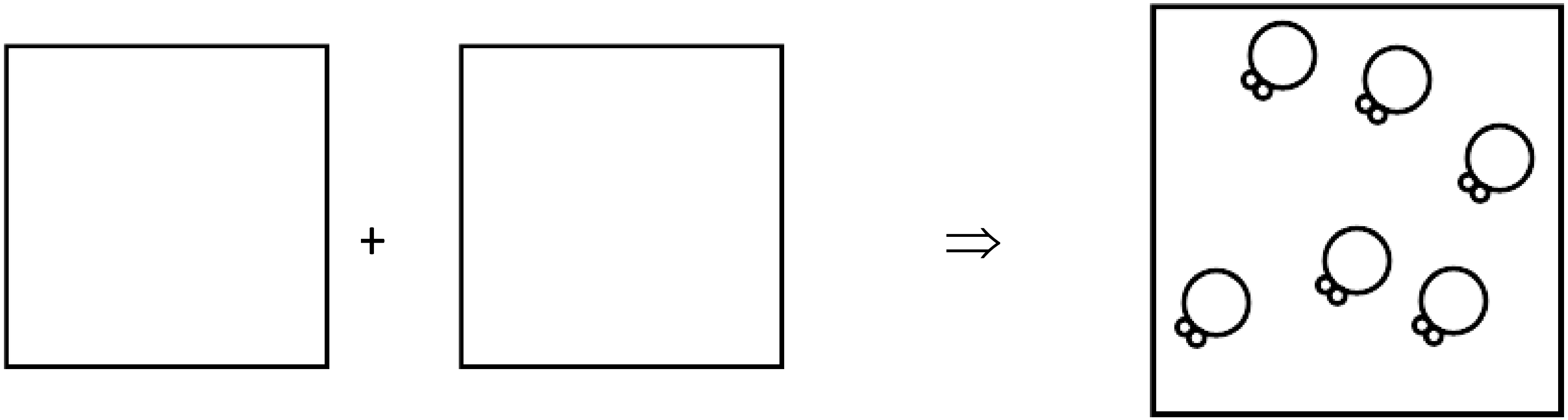


I.E: Atoms are conserved in physical and chemical processes.

- Chemical reaction is a simple rearrangement of atoms.
- No atoms are lost of gained.
- Particulate drawing are important part of this concept.

Draw a picture of the original reactions of the reaction (Proportions)

- $2H_2 + O_2 \Rightarrow 2H_2O$



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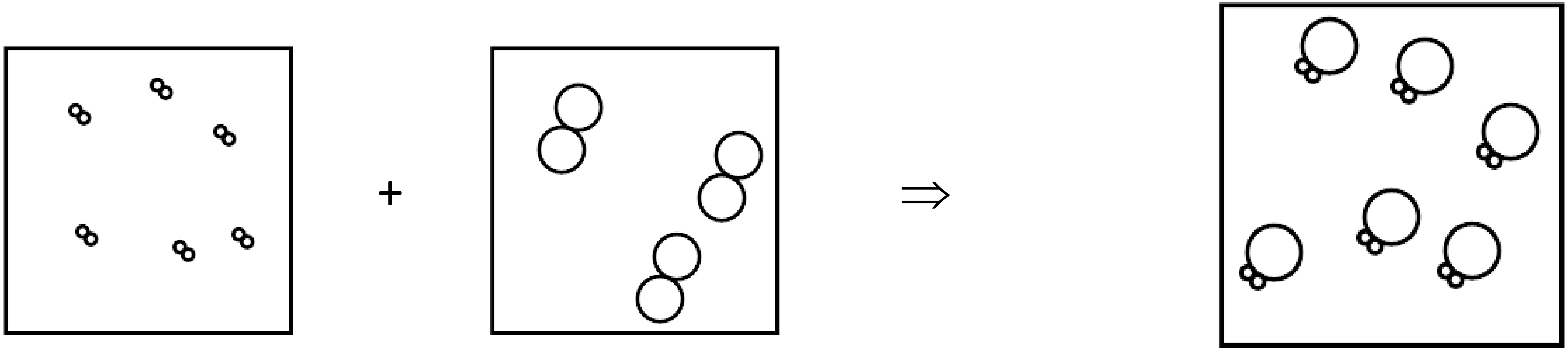


Gravimetric analysis: application of conservation of mass. (proportions)

- An unknown [M?] of 10mL Silver nitrate has a solution of NaCl added to it causing the silver Chloride to enter the solution. If 20mL of .1 M NaCl is required to complete the precipitation, what was the concentration of the original $AgNO_{3.}$

- Reaction will likely be given but,

$$AgNO_3 + NaCl \Rightarrow NaNO_3(aq) + AgCl(s)$$

- Note the stoichiometric ratios

Gravimetric analysis -- Answer

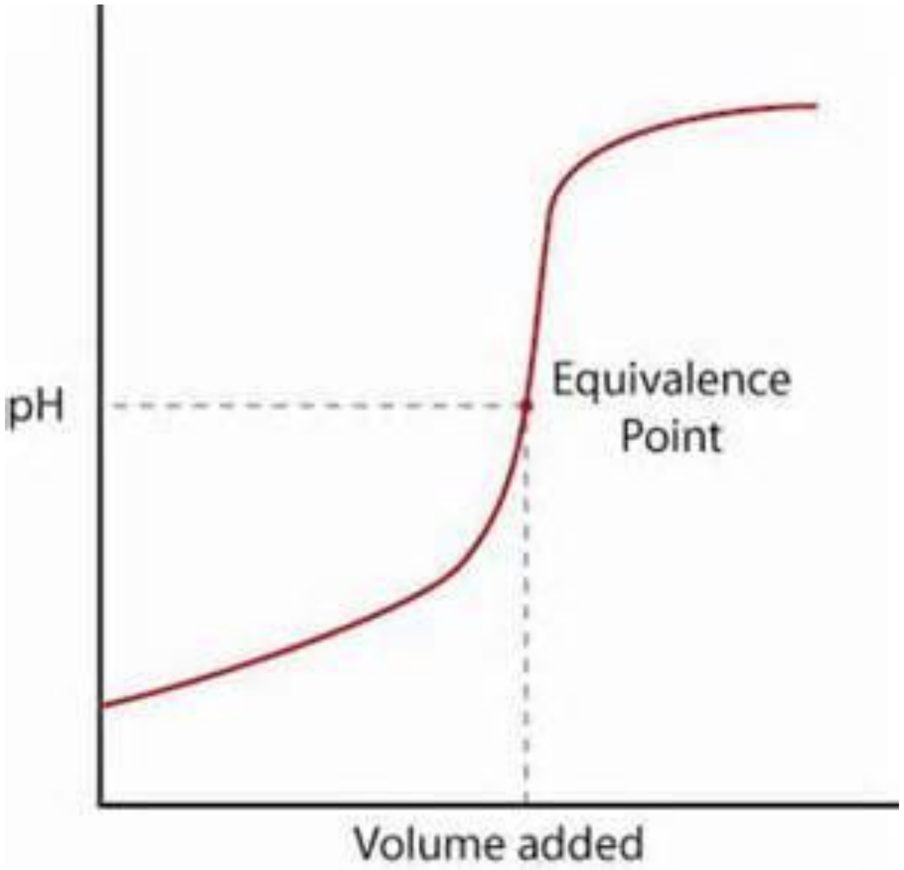
- What do we want?
- M = moles/L

Need:

- L = given at 10mL
- Moles: Determine the moles of NaCl use stoich ratio (1:1)
  - Moles = .1 * .02L = .002moles = moles of NaCl = $AgNO_3$
- M = moles/L .002/.01 = .2M
- NOTE: This problem did not require a calculator and likely could have been solved with simple proportions. Twice the volume would yield a concentration that was twice!

Titration proportions

	Known	UnKnown
M		
mol		
L		



Empirical vs. Molecular formulas

- Atoms come in whole number and therefore molecules have whole number of atoms. (Daltons law of definite and multiple proportions)
- Empirical formula: Simplest ratio of atoms
  - All ionic compounds are empirical
    - NaCl is the simplest unit of the crystal sometimes referred to as the formula unit.
  - Molecules have a specific number of atoms in a molecule.
  - Example: CO_2 1:2 ratio
  - Example: Oxalate ($C_2O_4^{-2}$)
    - Both of these examples have the exact same ratio of atoms 1 Carbon: 2 Oxygens

1. Complete combustion of a sample of a hydrocarbon in excess oxygen produces equimolar quantities of carbon dioxide and water. Which of the following could be the molecular formula of the compound?

   (A) C_2H_2
   (B) C_2H_6
   (C) C_4H_8
   (D) C_6H_6

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Notice: You are given equal quantities

Moles CO_2 = moles H_2O

Since 1 C/CO_2 and 2H/H_2O

Ratio 1C: 2H

Answer is C

#4

Ion	Ionic Radius (pm)
Zn^{2+}	74
Ca^{2+}	100
Ba^{2+}	135

$Zn(s)$ is used to reduce other compounds in chemical reactions. If a chemist needs a substance that is more effective in its reducing ability, which of the following species would be the best choice?

(A) Na

(B) H^+

(C) K^+

(D) Cl^-

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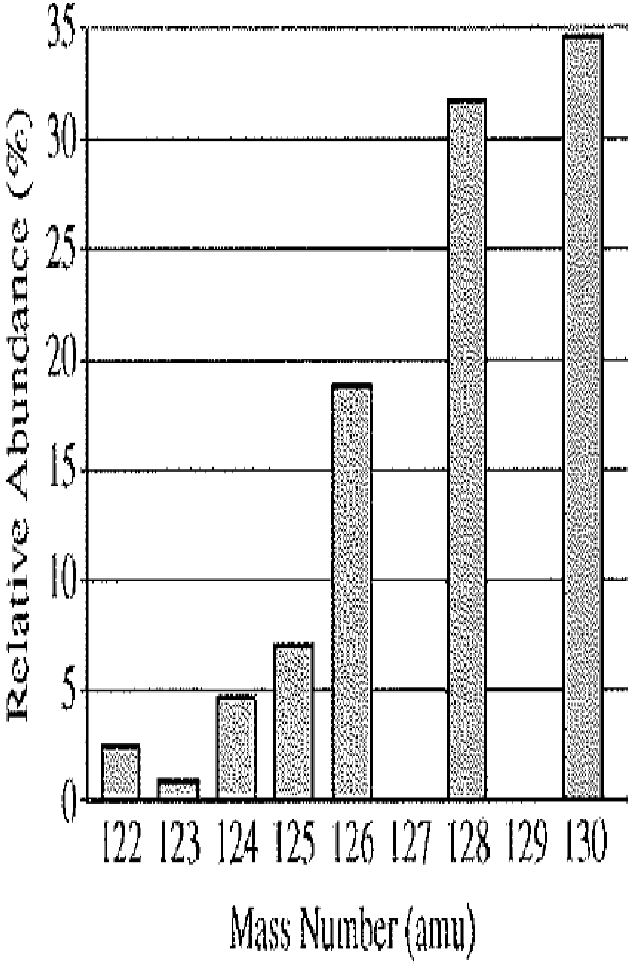
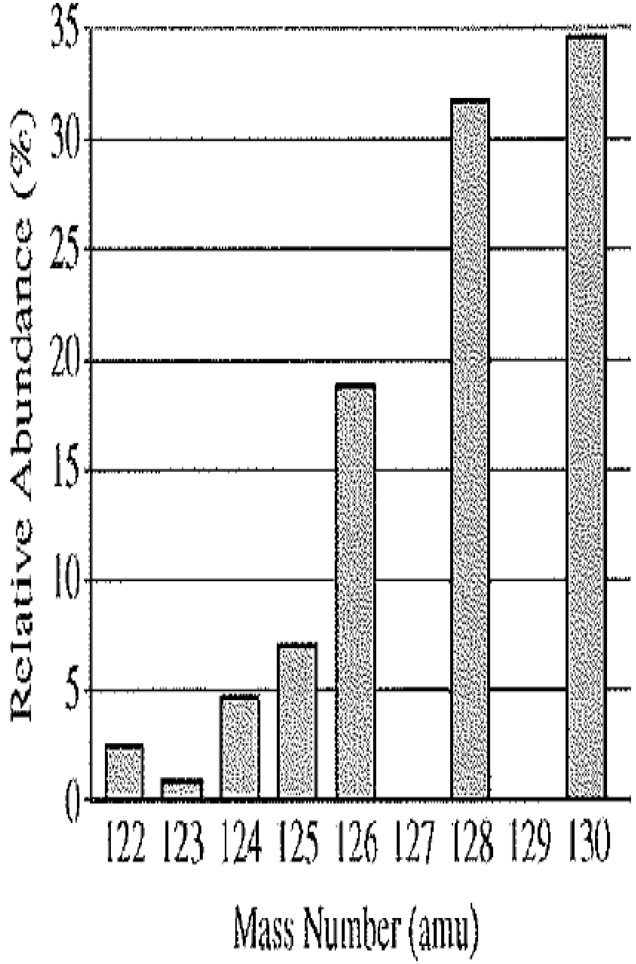
Think under pressure! Zinc is a metal that
Will do what metal do! Oxidize to a cation.
$Zn = Zn^{2+} + 2e^-$

What from this list will also do this?
Answer is A



10. The elements I and Te have similar average atomic masses. A sample that was believed to be a mixture of I and Te was run through a mass spectrometer, resulting in the data above. All of the following statements are true. Which one would be the best basis for concluding that the sample was pure Te?

(A) Te forms ions with a –2 charge, whereas I forms ions with a –1 charge.
(B) Te is more abundant than I in the universe.
(C) I consists of only one naturally occurring isotope with 74 neutrons, whereas Te has more than one isotope.
(D) I has a higher first ionization energy than Te does.



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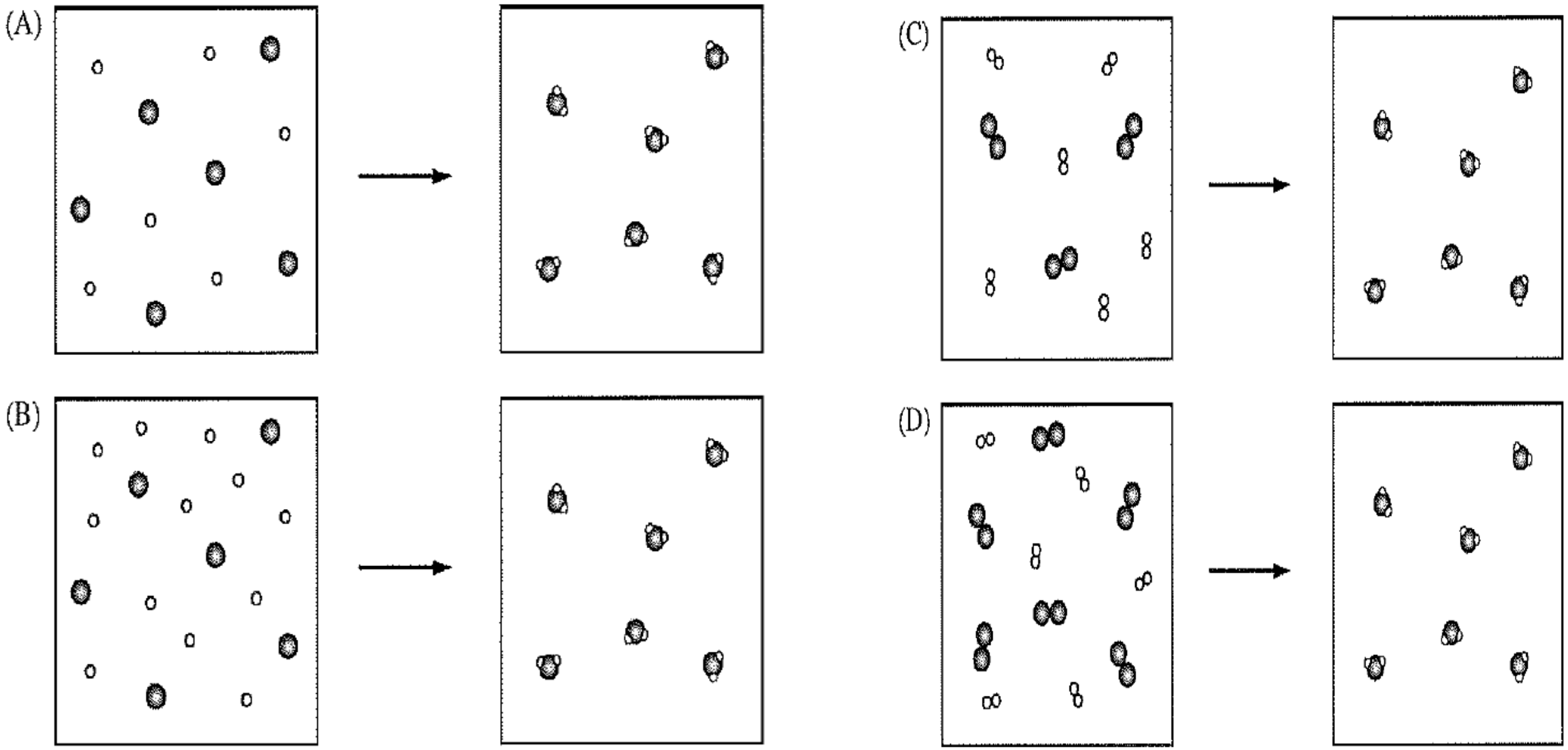
Electrons have no mass So any
Reference to electrons will have no
Baring on this topic. 127 is missing
This is the iodine average.

What is the reaction represented by combustion of hydrogen and water?



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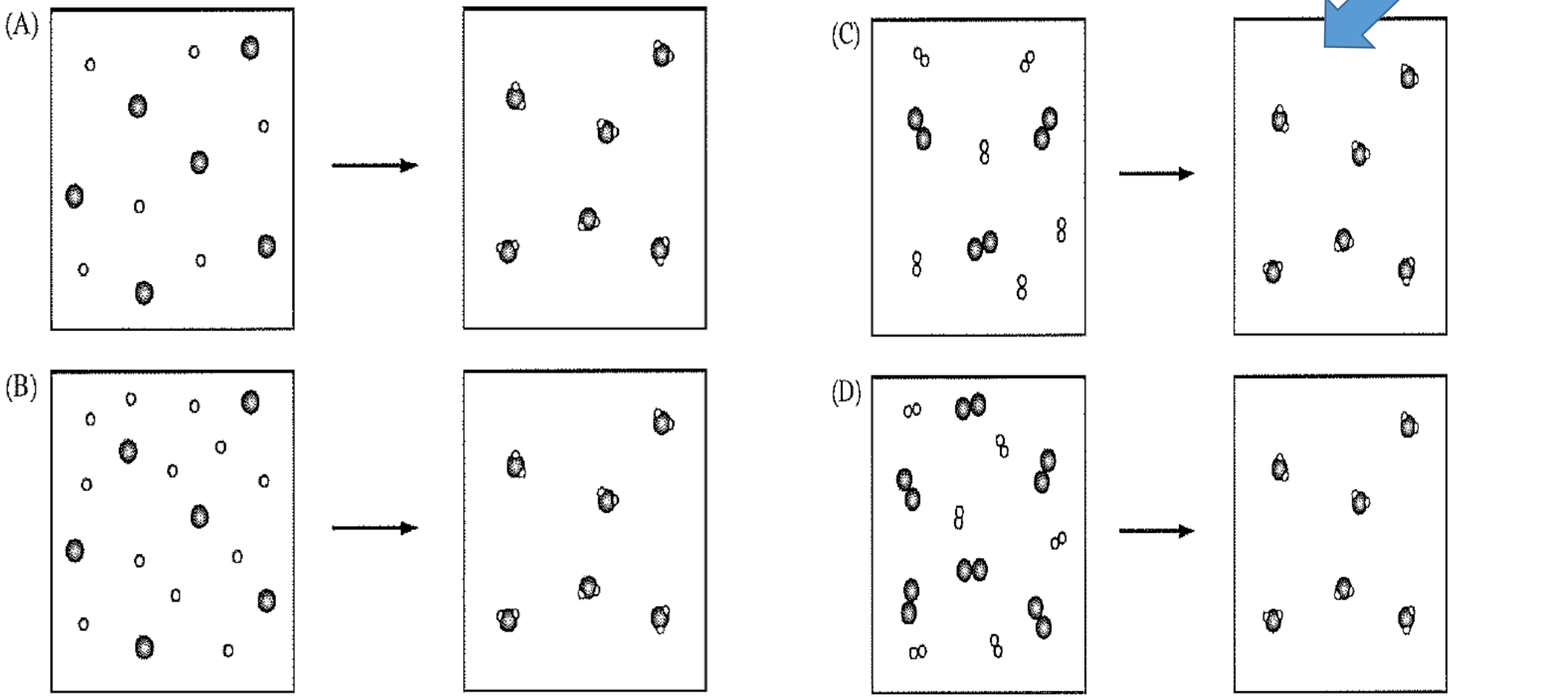

21. The mass percent of carbon in pure glucose, $C_6H_{12}O_6$, is 40.0 percent. A chemist analyzes an impure sample of glucose and determines that the mass percent of carbon is 38.2 percent. Which of the following impurities could account for the low mass percent of carbon in the sample?

(A) Water, H_2O
(B) Ribose, $C_5H_{10}O_5$
(C) Fructose, $C_6H_{12}O_6$, an isomer of glucose
(D) Sucrose, $C_{12}H_{22}O_{11}$

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Read Closely… The chemist is analyzing for C. Any chemical that does not contain C is out. There is nothing wrong with the chemists Technique. Which substance below would have a % Carbon that is lower and would then throw off the value.

$C_{12}H_{22}O_{11}$ is the same as glucose except a water Molecule is removed. So its % is higher. If you don't Know this you will have to sketch out the math.

Answer is B.

40. Silicon crystals are semiconductors. Which of the following is a correct reason for the increase in the conductivity of Si crystals when a small fraction of Si atoms are replaced with those of a different element?

(A) P atoms introduce additional mobile negative charges.
(B) P atoms introduce additional mobile positive charges.
(C) Ge atoms have more electrons than Si atoms have.
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It recognizes that phosphorus atoms contain one additional valence electron over the silicon atoms that make up the bulk of the material, and that these additional mobile charges will increase the conductivity of the silicon material. A is correct answer

41. N_2 molecules absorb ultraviolet light but not visible light. I_2 molecules absorb both visible and ultraviolet light. Which of the following statements explains the observations?

(A) More energy is required to make N_2 molecules vibrate than is required to make I_2 molecules vibrate.

(B) More energy is required to remove an electron from an I_2 molecule than is required to remove an electron from a N_2 molecule.

(C) Visible light does not produce transitions between electronic energy levels in the N_2 molecule but does produce transitions in the I_2 molecule.

(D) The molecular mass of I_2 is greater than the molecular mass of N_2.

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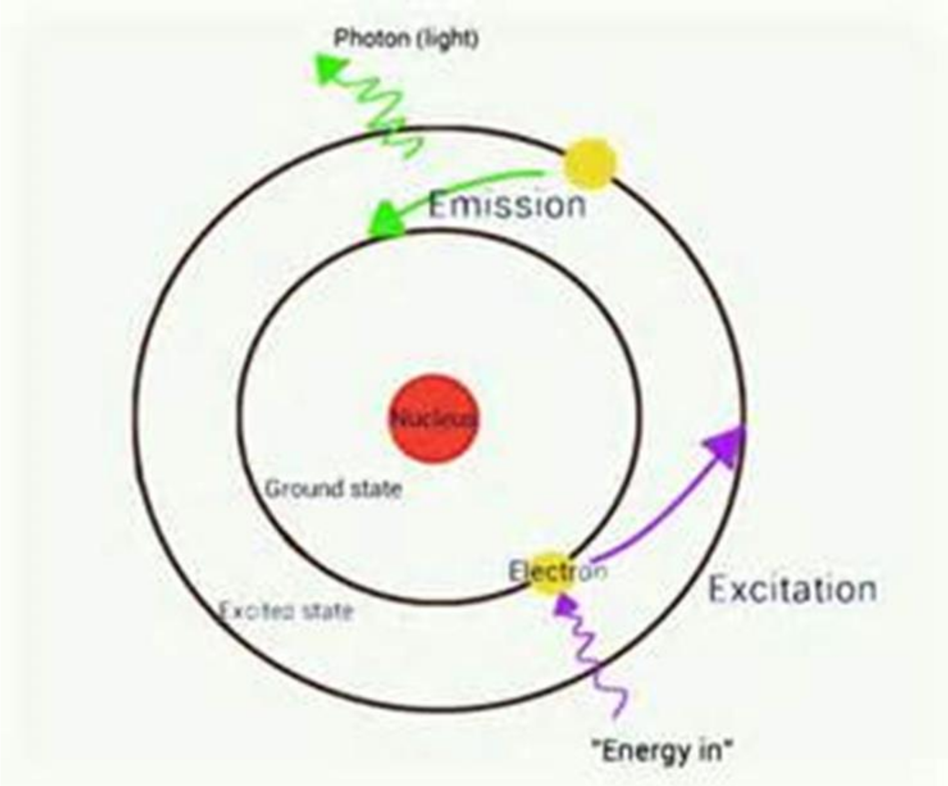
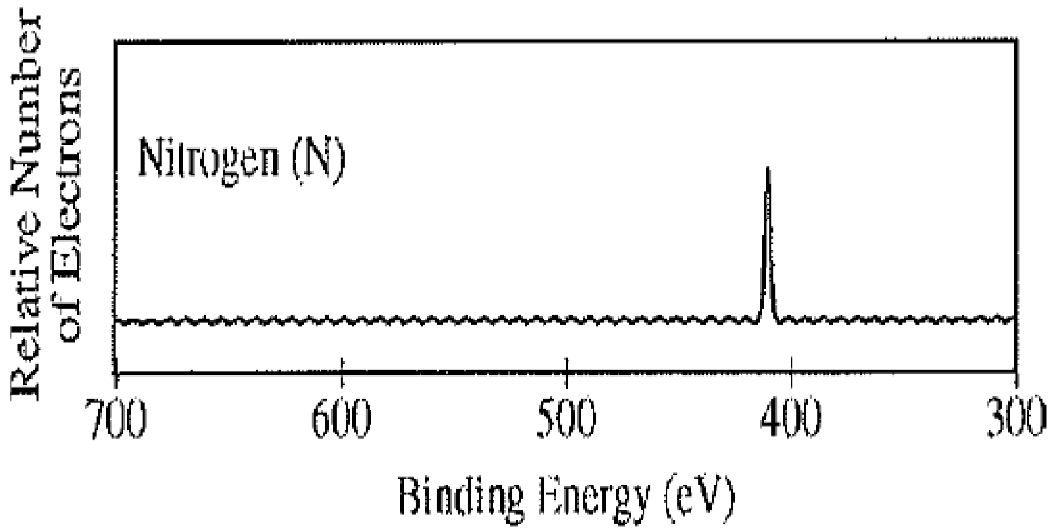
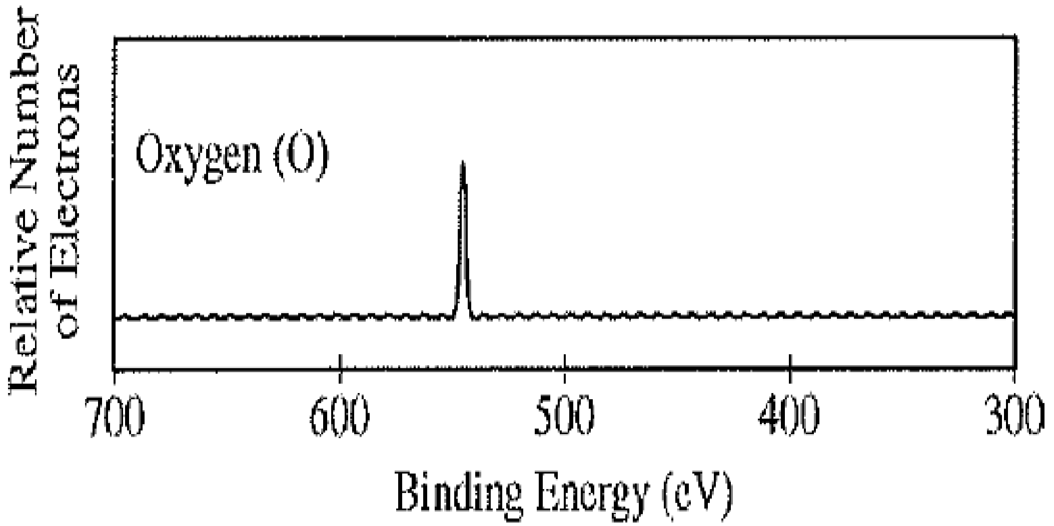






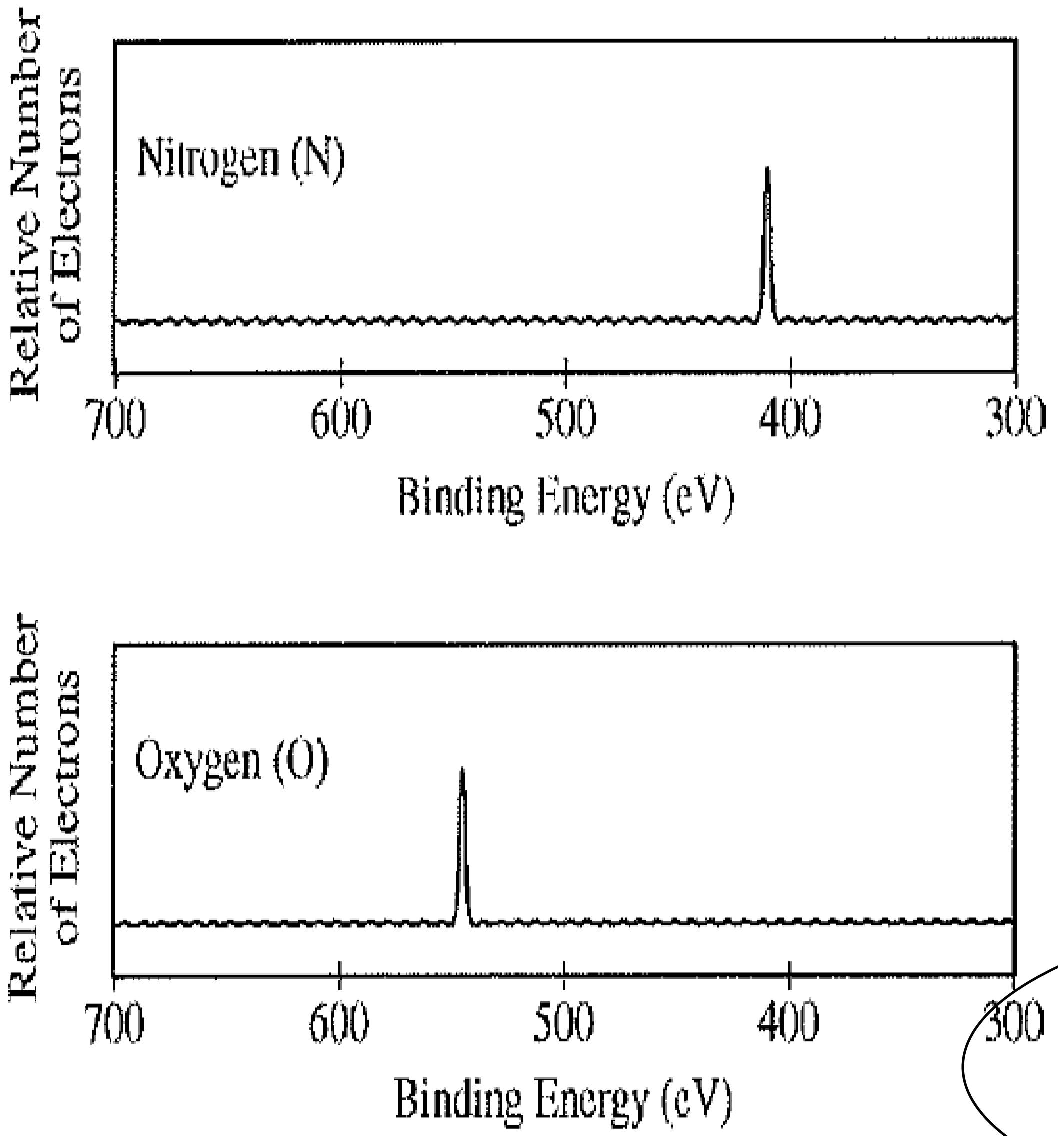
43. The photoelectron spectra above show the energy required to remove a 1*s* electron from a nitrogen atom and from an oxygen atom. Which of the following statements best accounts for the peak in the upper spectrum being to the right of the peak in the lower spectrum?

(A) Nitrogen atoms have a half-filled *p* subshell.

(B) There are more electron-electron repulsions in oxygen atoms than in nitrogen atoms.

(C) Electrons in the *p* subshell of oxygen atoms provide more shielding than electrons in the *p* subshell of nitrogen atoms.

(D) Nitrogen atoms have a smaller nuclear charge than oxygen atoms.



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