Unit 2

Chapter 4-Atoms and Elements, continued

I Ions

II ISOTOPES-Tools

A. Tools

1. Atomic number, ℤ,, equals the number of protons

2. Mass number, 𝔸 , equals the sum of protons and neutron (nucleons)

3. Electrons, e-, equals the number of protons in a neutral atom

4. Subatomic Particles = e^- , p^+ , n

Summary:

Particle	Symbol	Charge	Mass	Location

Remember: opposites attract

B. Atomic Isotopes - Atoms with the same atomic number (# p^+), but different number of neutrons
→ Different mass
Examples:

Isotopic Symbol	Mass #, A	Atomic #, Z	# p^+	# n	% natural abundance
N-14					
N-12					
N-13					

Atomic Isotopes - Problems:

1. Give the isotopic symbol for an atom of Be that has the same number of p^+, n, and e^-

2. Find the number of protons, neutrons and electrons in Ne-22

C. IONS-Ions are atoms with a charge

IONS

ANIONS

CATIONS

Isotope problems

1. a. Calculate the number of protons, electrons and neutrons of an N-14 atom that has a -3 charge.

b. How many subatomic particles does this ion have?

2. An isotope of copper that has 3 more neutrons than copper-63

3. An atom of chromium which ahs the same number of electrons, protons, and neutrons.

4. An atom with 5 more neutrons and 2 more protons than ^{31}P.

5. An isotope of sulfur that contains the same number of neutrons as silicon-28.

6. An isotope of iron that contains the same number of subatomic particles as nickel-59.

Chapter 5-Molecules and Compounds

CHEMICAL FORMULAS – HOW TO REPRESENT COMPOUNDS

I. COMPOUNDS - Two or more elements chemically combined in definite proportions.

COMPOUNDS

IONIC COMPOUNDS	MOLECULAR COMPOUNDS
Metal - Nonmetal	Nonmetal-Nonmetal

II. Ionic compounds

Formation of ions:

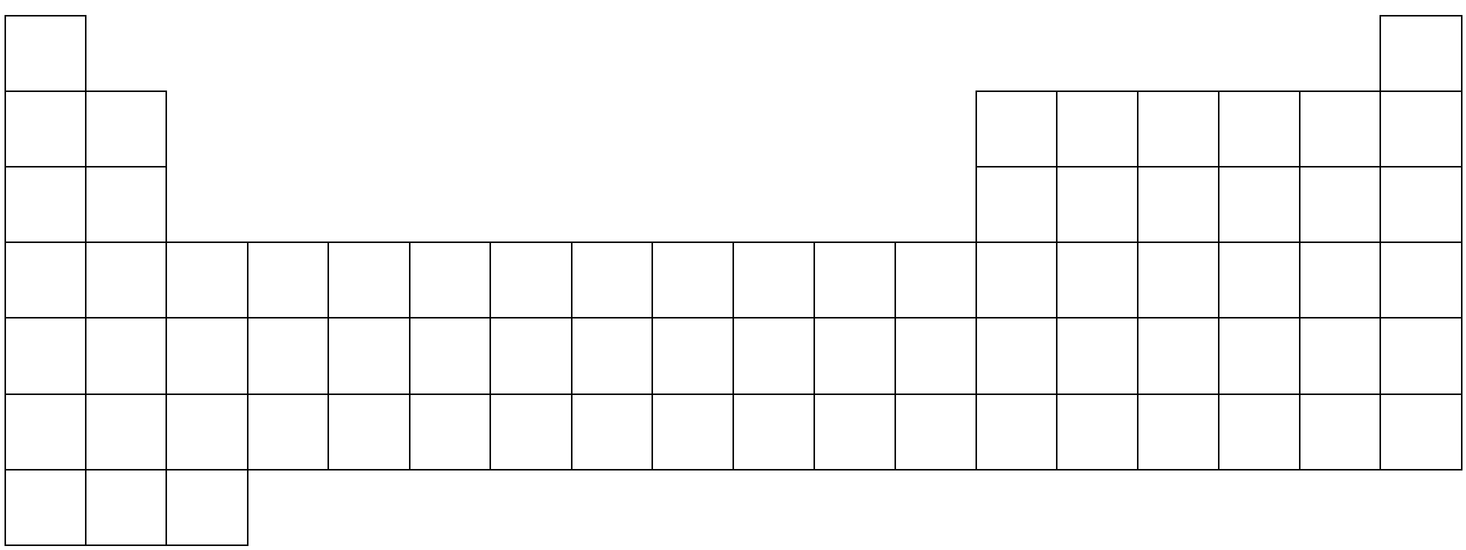
Metals form cations (+) ions:

Nonmetals form Anions (-) ions:

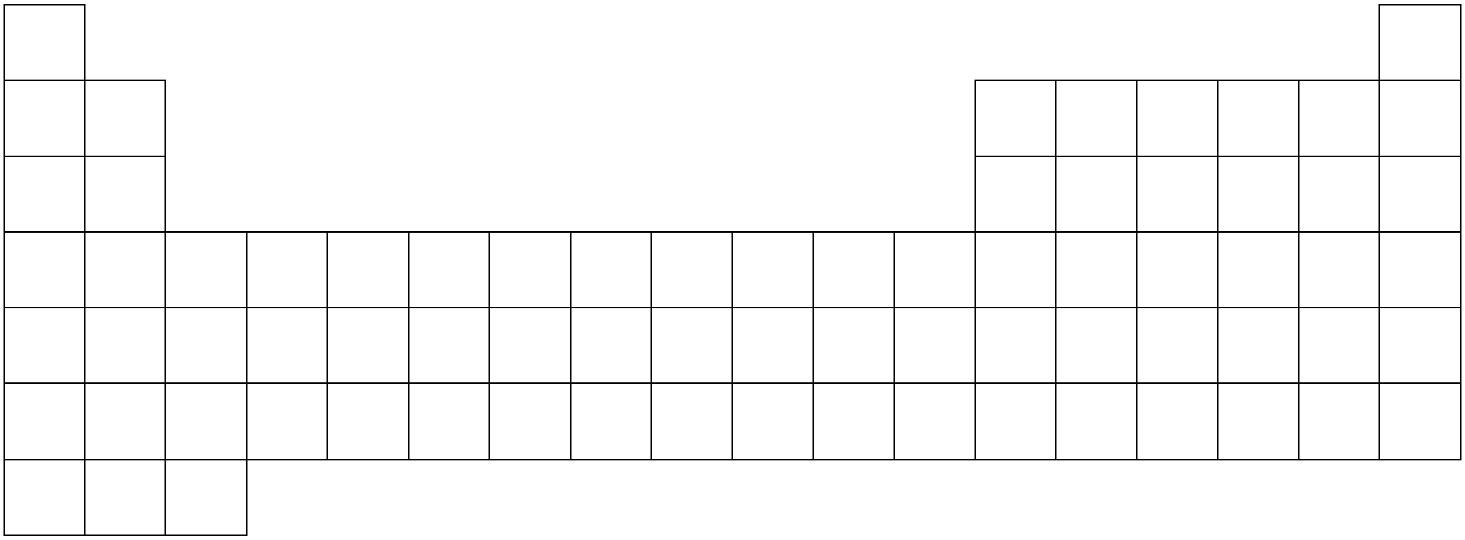
III. Cation and Anions

A. Metallic Cations - (+ charge)

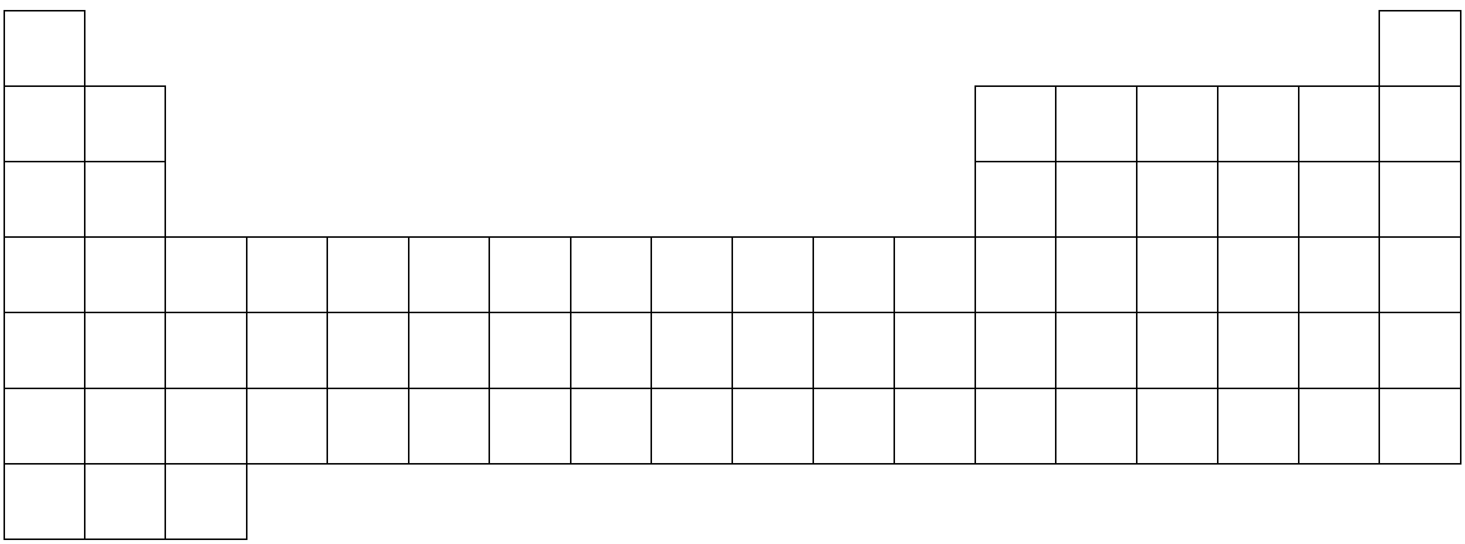
1. Fixed Charged cations are those metals that form only one type of ion.



2. <u>Variable charged cations</u> are those metals that form <u>more</u> than one type of ion.



B. Nonmetal Anions (-) charge



IV. Naming Ionic Compounds

1. Fixed metal → name as is

2. Variable charged metal- remember the (roman numeral) or use the classical "common" name.

V. Polyatomic Ions

A group of atoms bonded together to form an ion

N		S		P	
NO_2^{1-}	Nitrite	SO_3^{2-}	Sulfite	PO_3^{3-}	Phosphite
NO_3^{1-}	Nitrate	SO_4^{2-}	Sulfate	PO_4^{3-}	Phosphate

VI. Naming compounds with polyatomic ions

VII. Chemical Formulas

Key: Compounds are neutral → no net charge

VIII. Molecular compounds

Nonmetal - Nonmetal

Variable combinations

Ex.

1. Know prefixes:

Prefix	Number
Mono-	1
Di-	2
Tri-	3
Tetra-	4
Penta-	5
Hexa-	6
Hepta-	7
Octa-	8
Ennea- /Nona-	9
Deca-	10

2. Naming formula:

Prefix element #1 + prefix stem of element #2 + ide

Ex.

IX. HYDRATES

Ionic compounds that incorporates H_2O in their crystalline structure (water of hydration)

ex. $CuSO_4 \cdot 5\ H_2O$

X. ACIDS

Formula starts with a "H" + (aq)
[H_2O is excluded]

Ex. HCl (aq) "Dissolved in water" → The HCl **must** be in H_2O to have the properties of an acid.

ACIDS

Binary Acid	Oxyacid/ Ternary Acid
Does not contain "O"	Contains "O"

A. Binary Acids (no "O")

Naming: Hydro + stem of element + ic Acid

Ex.

Exception: H_2S-->

B. OXYACIDS (contains "O")

Naming Formula:

Ion name But Change ite--> ous / ate--> ic + Acid

KEY: Recognize the ion part of the Acid

ACID	ION

NOMENCLATURE

ACID

Formula starts with an "H" + (aq)
(Water is excluded)

ACID

oxyacid: contains oxygen

binary acid: does not contain oxygen

EXAMPLES:

oxyacid	binary acid
HClO	HCl
hypochlorous acid	hydrochloric acid

IONIC COMPOUND

Metal-Nonmetal

IONIC COMPOUND

Fixed charged metal: Group IA,IIA, Al, Ga, Cd, Zn, or Ag
--> name as is

KBr
Potassium bromide

Variable charged metal: All other metals
-->place charge in ( ) as roman numerials/ know classical names

CuOH
Copper (I) hydroxide

MOLECULAR COMPOUND

Nonmetal-Nonmetal

MOLECULAR COMPOUND

When naming, use prefixes (mono, di tri...etc.)

CO
Carbon monoxide

HYDRATES: CONTAINS H_2O in the chemical formula - use prefix + Hydrate

$CuSO_4 \cdot 5H_2O$ -Copper (II)) sulfate pentahydrate

Nomenclature Examples

Chapter 6-CHEMICAL COMPOSITION

CHEMICAL FORMULA CALCULATIONS

I. FORMULA WEIGHTS = $\sum$ Mass of all atoms

Atomic level

one formula unit of Al_2O_3 =

II. THE MOLE

1 mole of particles = 6.02×10^{23} particles

Avogadro's number → memorize!!

Conversions

$$\frac{\text{1 mole H atoms}}{6.02 \times 10^{23}\text{ atoms}} \quad \text{or} \quad \frac{6.02 \times 10^{23}\text{ H atoms}}{\text{1 mole atoms}}$$

Problem: How many Cu atoms in 6.0 mol Cu?

Know: 1 mol Cu = 6.02×10^{23} atom

III. MOLAR MASS (molecular wt.)

1 mole = AMU weight numerically in grams

26
Fe
55.85

Atomic wt.
55.85 AMU
{1 atom}

Molar mass
55.85 g
= 1mole of Fe atoms
= 6.02×10^{23} Fe atoms

Conversion factors:

Problems:

1a. How many moles of Fe in 33.0 g of Fe?

b. How many atoms is this?

IV. MOLECULAR COMPOUNDS AND IONIC COMPOUNDS

V. MOLES AND CHEMICAL FORMULAS

N_2O_5

2 atoms N
5 atoms O
= 1 molecule N_2O_5

2 mole N
5 moles O
= 1 mole of N_2O_5

Ratios:

Problems:

1. How many moles of N in 13.5 moles of N_2O_5?

2. How many moles of O in 13.5 moles of N_2O_5?

VI. MOLES AND CHEMICAL CALCULATIONS:

subscripts in chem. formula

#atoms/molecules, Formula Units ⟷ mole_1 ⟷ mole_2 ⟷ Mass,grams

$6.02 \times 10^{23}\ \frac{\textit{units}}{\textit{mole}}$ MM, *molar mass*

Problems:

1. How many atoms in 13.4 g of S?

2. How many K atoms in 3.0 g K_3P?

3. 1.50×10^{25} atoms weighed 398.7 g
   a. What is it's molar mass?

   b. What element is this?

4. What is the **total** number of atoms in 0.20 grams of K_3P?

5. What is the mass in mg of 1 atom of Al?

6. How many atoms of O are in 32 kg of phosphoric acid?

7. How many grams of octane C_8H_{18} contain 3.02×10^{21} atoms of carbon?

8. How many molecules of water contain 123.44 grams of oxygen

9. How many atoms of oxygen are in a sample of glucose $C_6H_{12}O_6$ that contains 43.55 grams of hydrogen?

10. How many grams of glucose contain 5.226×10^{24} atoms of hydrogen

11. How many total atoms are in a sample of 3.09×10^{25} molecules of glucose.

VII. PERCENTAGE COMPOSITION

$$\% \text{ BY WT.} = \left(\frac{\text{WT. of Element}}{\text{Total mass of compound/sample}} \right) \times (100)$$

What is the % composition of NaCl?

VIII. EMPIRICAL FORMULA

Empirical formula shows the smallest ratio of atoms in a compound.
Examples:

IX. Calculation of Empirical Formula

Experiment: Elemental Analysis gives %wt of elements

Problem: Find the empirical formula for a compound containing: 11.19 %H and 88.89% O

Step 1. Express the percent in grams → Assume 100 g of material.
IF THE DATA IS ALREADY IN GRAMS GO TO STEP 2!!!

Step 2. Change the grams into moles

Step 3. Change the numbers to whole numbers by dividing by the smallest number.

Problem #2 A 10.00 g sample was analyzed as: 5.293 g Al and the rest is Oxygen. What is its empirical formula?

Multipliers when the simplest mole ratio is not a whole number:

$Al_1O_{1.5}$ $C_1H_{1.25}$ $C_1H_{1.33}$ $C_1H_{1.66}$

X. Calculation of Molecular Formula

Problem: A compound contains 38.7 % C, 9.7% H, and 51.6 %O. The molar mass of the compound is 62.07 g/mol

STEP 1. Calculate the Empirical Formula

STEP. 2 Calculate the Empirical Formula weight.

STEP. 3 Determine the number of E.F. units in the molecular formula
{Divide the molar mass by the E.F. wt.}

Problem. A 5.00 gram sample contains 4.69 g C and the rest is Hydrogen. It's molar mass is 125 g/mol

Deeper PROBLEMS

1. A 2.00g sample of lithium metal is burned in oxygen atmosphere to produce 4.31 g of a lithium-oxygen compound. Determine the compound's empirical formula.

2. A 3.750 g sample of the compound responsible for the odor of cloves (containing only C, H, and O is burned in a combustion analysis apparatus. The mass of CO_2 is produce is 10.05 g and the mass of H2O produced is 2.470 g. What is the empirical formula of the compound?

3. By analysis, a compound with the formula KClOx is found to contain 28.9% chlorine by mass. What is the value of x?

4._A 7.503 g sample of metal is reacted with excess oxygen to yield 10.498 g of the oxide MO. Calc the molar mass of element M.

Chapter 9-ELECTRONS IN THE ATOM AND THE PERIODIC TABLE

I. Electronic Arrangement in Atoms

A. Background:

Where are the electrons? How do they move? Randomly? Set patterns?
Expt:

Atomic Model- Each electron has a fixed, specific energy due to the distance from the nucleus.

B. Quantum-Mechanical Orbitals - Electron structure-Section 9.6

Background:

Electrons in the atoms are found in **principal energy levels (n)** also called shells.

1. In each principal energy level,n, e- move within orbitals.
Principal energy levels are divided into **sublevels** (or subshells), which consists of **orbitals.**
   a. Shapes of orbitals

   b. There are 2 e^- per orbital

   c. The maximum number of each type of orbitals allowed per Principal Energy level (n) and number of electrons in each sublevel (orbital set) are as follows:

ORBITAL	# of orbitals Per Energy Level	*X 2 e^-/orbital*	TOTAL # of e^-

d. Relative Orbital energy.......

(1) within each n value (principal energy level)

(2) for the same type of orbital but different energy levels

(3) within a p "set" of orbitals {same sublevel}

2. ELECTRON CONFIGURATION

The filling order of electrons

1. e^- will occupy the lowest energy orbital possible

2. Orbital Designation

3. Using the periodic table......
   a] Period = n for the s & p sublevels

   Period = n-1 for the d sublevel.

   b] Sublevel {s, p, and d} areas in the periodic table.

   c] Obtain the e^- configuration by going across periods in increasing atomic number (increasing ene

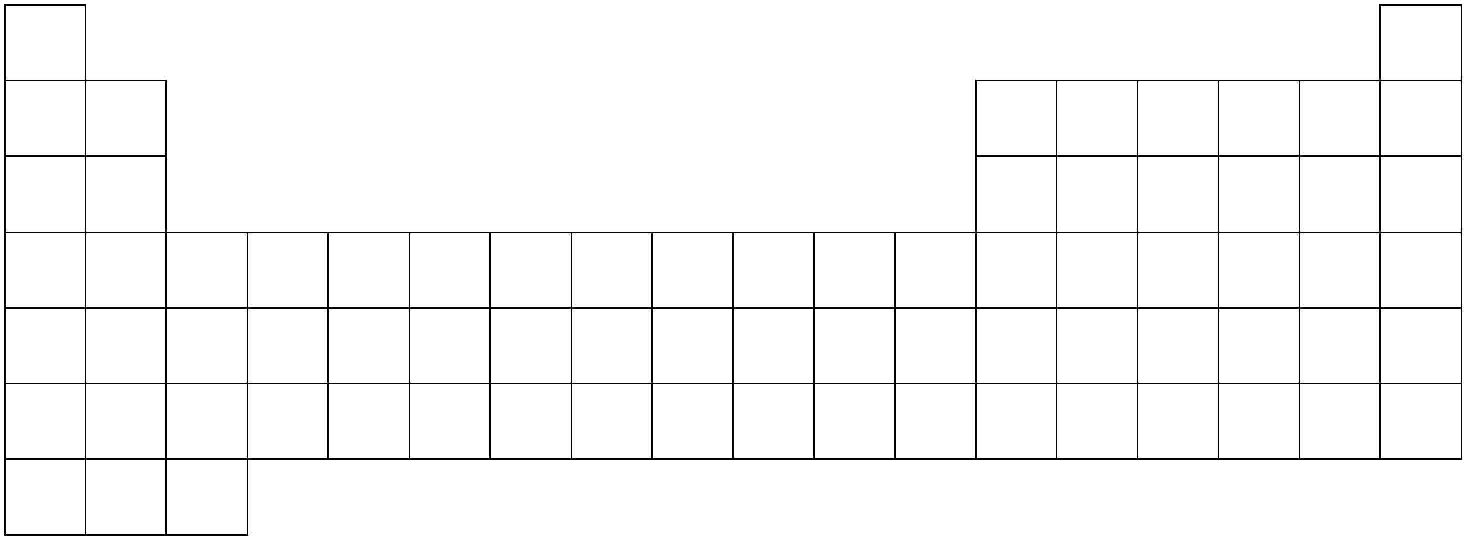
   (1) Each square in the periodic table = 1 e-

   (2) Period = Energy level, n, for s and p

1s 2s 2p 3s 3p 4s 3d 4p 5s 4d 5p.........

3. Orbital Diagrams

4. Electron configurations and Orbital diagrams



4. Electron configurations and Orbital diagrams

Examples:

1. O

2. Si

3. Br

4. Na^+

5. S^{2-}

5. Ne

6. Mg

d. CORE NOTATION

Ca = [Ar]$4s^2$

↰ the "core" must be the previous inert gas

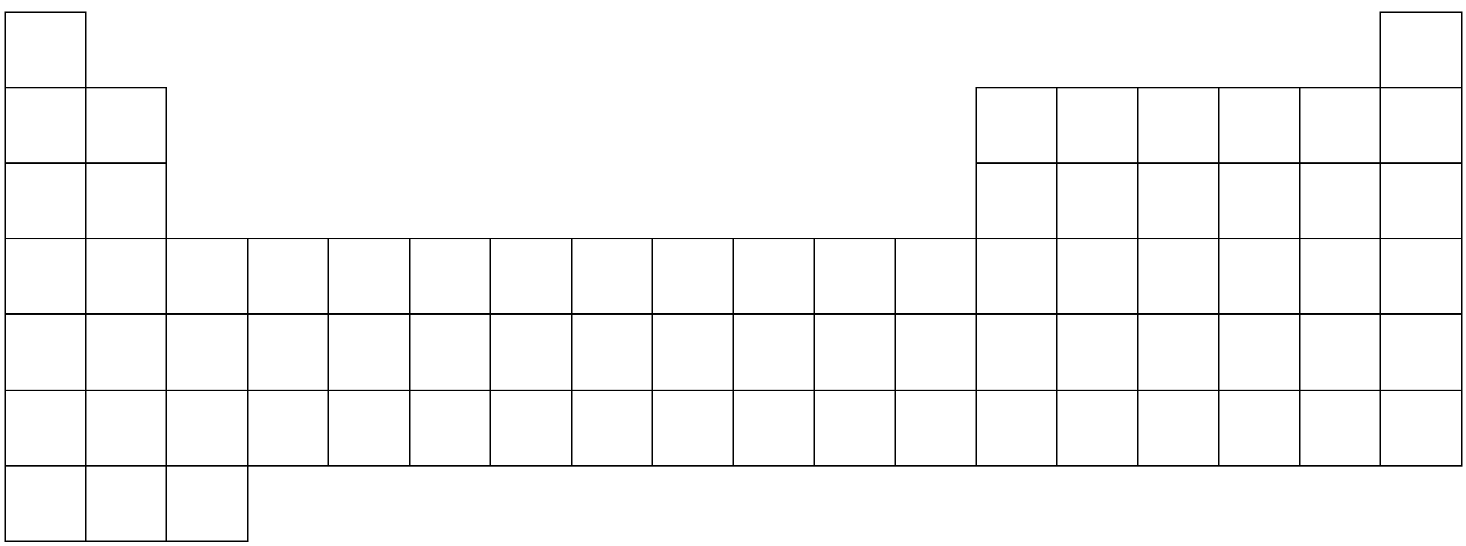
Examples:

<u>II. Valence electrons</u> (High energy electrons)

The electrons in the outermost shell (energy level). Valence electrons are involved in reactions.

(Rem: # valence e^- = the group number for the "A" subgroup elements)

ex.



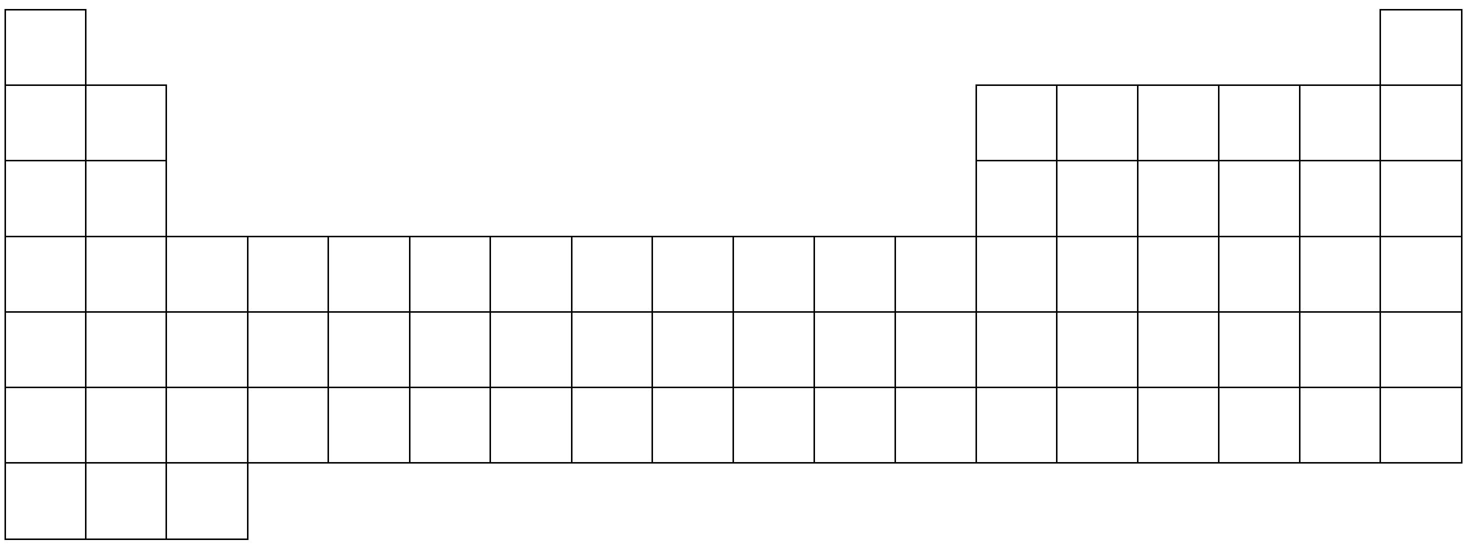
III. Periodic Trends

The periodic table can be used to predict certain characteristics of element. These predictions are based upon certain periodic trends.

A. Atomic Size [Atom Radius]

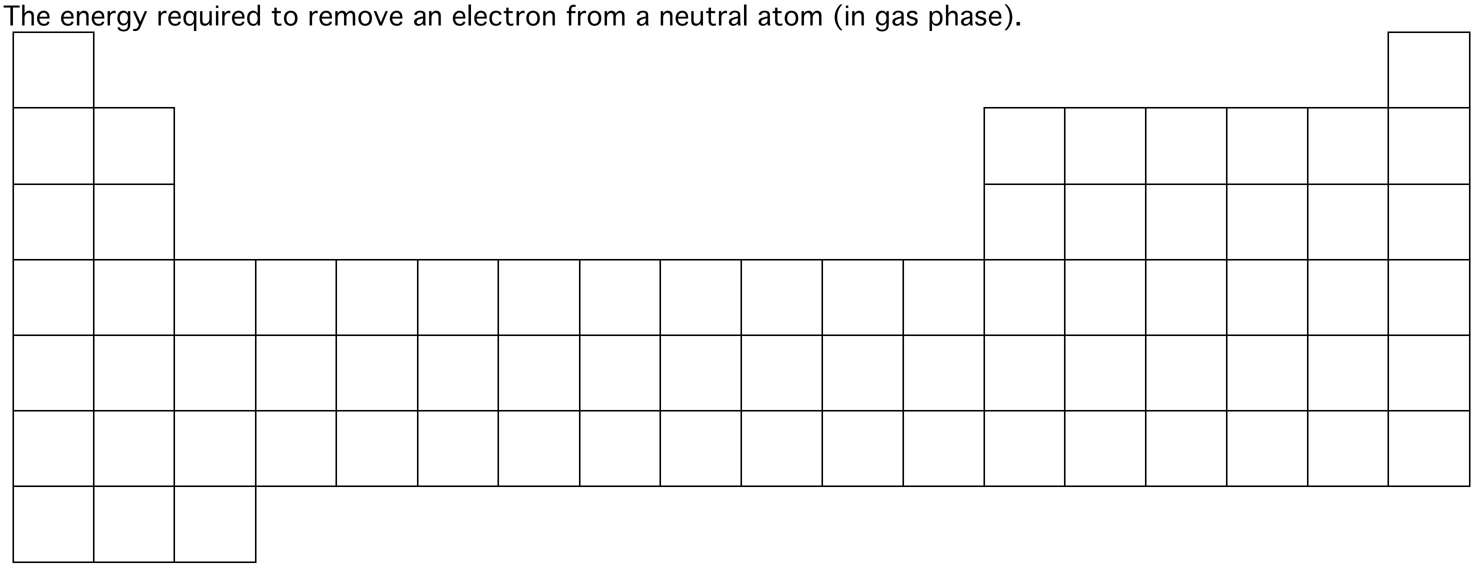
1. As the number of shells increases, the radius size increases

2. As you go left to right across a period, the radius size decreases



B Ionization Energy

The energy required to remove an electron from a neutral atom (in gas phase).

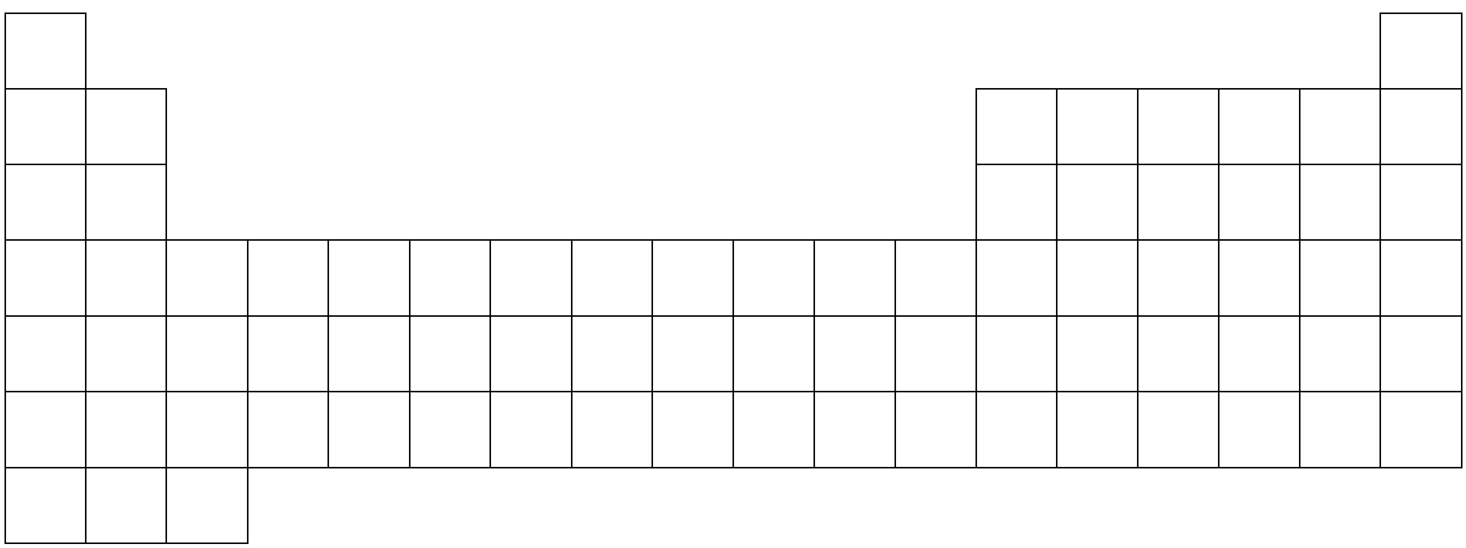


c. Metallic Character The amount of energy released or absorbed when an electron is added to an atom to form a (-) ion [anion], in gas phase.

1. As the number of shells increases, metallic character increases

2. As you go left to right across a period, metallic character decreases

d. Electron Affinity The amount of energy released or absorbed when an electron is added to an atom to form a (-) ion [anion], in gas phase.



Chapter 10-CHEMICAL BONDING

CHEMICAL BONDING→ The attractive interaction between two atoms or ions

I. Types: -

1. **Ionic Bond**- Cations (+ charged) and Anions (- c
   are held together by the
   attractive force of their (+) and (-) charges →Electrostatic force.

2. (Metallic Bonds)

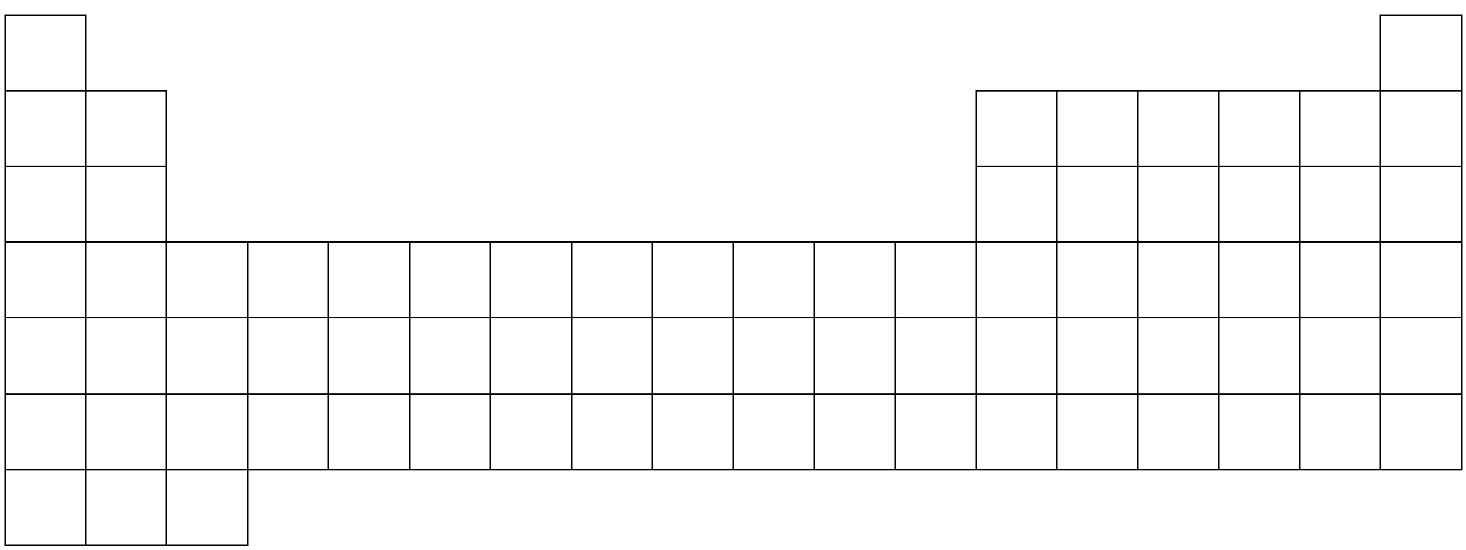
3. (Macro molecular crystals)

4. **Covalent Bonds**- Results from the sharing of a pair of electrons between two atoms.
   ex. CO H_2O
   a. Diatomic -

   b. Polyatomic-

II. Lewis Electron Dot Structures

1. ▪ = 1 e^-
2. Shows the number of valence electrons (reacting electrons)
3. Arrange electrons......



III. Octet Rule

Eight Valence electrons

Atoms react to obtain an "octet" (8 valence e^-)

Anions

Octet rule continued

Cations

VII. Covalent Bonds - A bond resulting from one or more shared electron pair(s)

Atoms share electron pairs to form an octet

Exception: Hydrogen forms a duet

VIII. Multiple Covalent Bonds

IX. Complex Electron Dot Structures

DRAWING ELECTRON DOT STRUCTURES

HOW TO:

1. Write e- dot structure for the individual atoms.

2. a) Add together the number of valence electrons for all the atoms
   (If it is an ion, you must add or subtract electrons accordingly)

   b) Divide the total number of e^- by 2: This will give you the number of e^- pairs available for bonding.

3. Determine which is the central atom
   a. The least represented atom that is **not H**
   b. Usually, the **first** atom in the chemical formula that is **not H**.

4. Arrange atoms symmetrically around the central atom.

5. Draw a single line (or 2 dots) between the central and outer atoms.

6. From the total number of valence electrons subtract 2 electron for each bond made.

7. Attempt to place the remaining electron pairs around the outer atoms to make an octet or duet (for H)

8. If an octet cannot be fulfilled; then, a double or triple bond must be formed.
   Warning: Do not use a double or triple bond unless you have to!

9. Each atom (except H) **must** have an octet!

HONC, a general rule (a help)

Examples:

Specific Electron Dot Cases:

a. Ions:

b. Oxy Acids

X. Electronetativity and Polarity: Why Oil and Water don't mix Sec. 10.8

Electronegativity- The ability of an element to attract electrons within a covalent bond

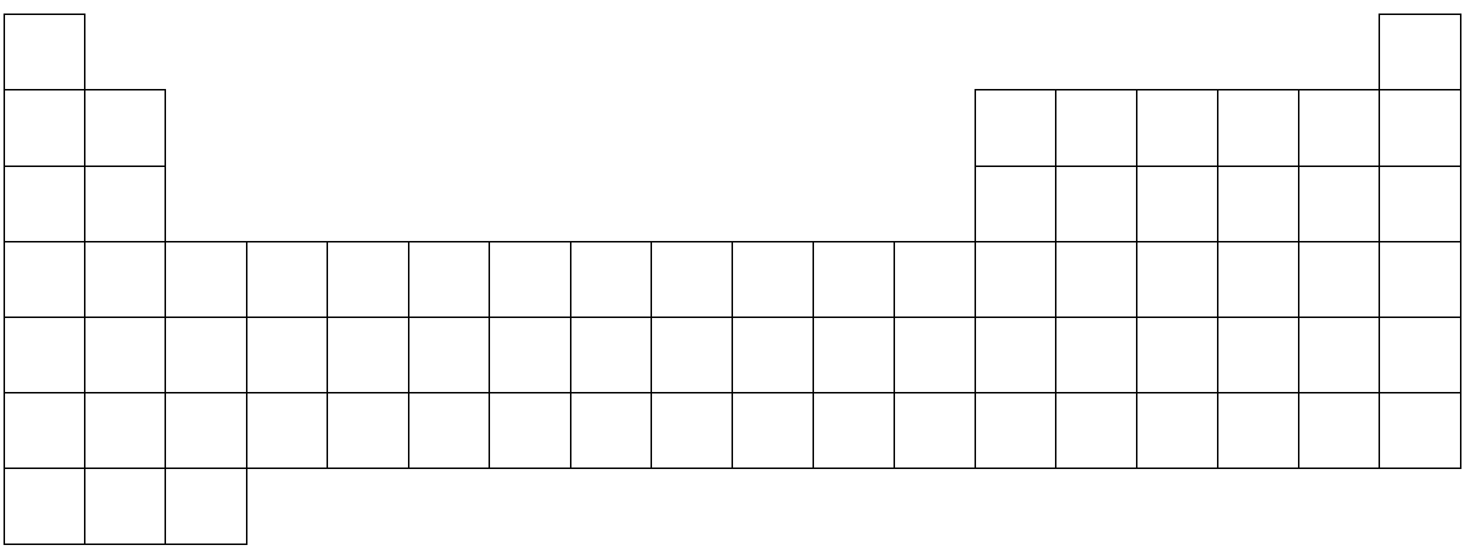
1. Bond Polarity.

Pure Covalent Bonding

Ionic bonding

Polar covalent bonding

3. Electronegtivity



4. <u>Molecular polarity</u> - net polarity of molecules

   Type 1-relative electronegativities

   a. Draw individual bond polarities, using relative electronegativity trends
   b. Find the net molecular polarity-using vector analysis by inspection.
   c. If there is a net polarity-the molecule is polar and has a **DIPOLE**!

   Type 2 – differences in electronegativity

   Δ EN 0 - .4 = covalent

   Δ EN .4 - 2.0 = polarcovalent

   Δ EN 2.0 + = ionic

Practice EXAM I A

SHOW ALL YOUR WORK. YOUR ANSWERS MUST HAVE THE CORRECT NUMBER OF SIGNIFICANT FIGURES AND UNITS. CORRECT SPELLING MUST BE USED.

1. COMPLETE THE FOLLOWING TABLE:

$^{A}_{Z}Sym$	Number of protons	Number of neutrons	Number of electrons	Mass number
______	16	_______	18	32
^{56}Fe	________	_______	_______	_______

2. Give the isotope symbol for the following:
   a. An anion of nitrogen with the same number of neutrons as oxygen-15

   b. An iron cation with the same charge and number of subatomic particles as $^{58}Co^{2+}$

3. Calculate the molar mass of Na_3PO_4

4. How many **grams of H** are there in 3.0×10^{25} molecules of H_2SO_4?

5. Name or give the chemical formula for the following:.

oxalic acid ________________________________	magnesium hydrogen carbonate ________________________________
mercurous nitride ________________________________	ammonium carbonate ________________________________
silver nitrate ________________________________	aurous iodide ________________________________
plumbic acetate ________________________________	iodine tribromide ________________________________
calcium peroxide ________________________________	hydrobromic acid ________________________________
potassium phosphide ________________________________	sulfurous acid ________________________________
nickelous permangante ________________________________	cobaltous sulfide ________________________________

CS_2

$Ni(NO_2)2$

Ba_3N_2

$Ca(OH)_2$

$Sr(HSO_3)_2$

$H_2CO_3(aq)$

SO_3

HF

$HC_2H_3O_2(aq)$

N_2O_3

HCN(aq)

KH

Co_2O_3

$Bi(NO_3)_3$

$HClO_3(aq)$

N_2O_5

$Hg(HCO_3)_2$

PbO_2

$HBrO_2(aq)$

Au_3PO_4

$Cu(Cl0)_2$

$Al(OH)_3$

6. The percentage composition of a compound is 63.133% C, 8.831% H, and 28.04% O.

   The Molar mass = 171.21 g/mol

   a. What is its empirical formula?

   b. What is its molecular formula?

7. The chemical formula of DDT is $C_{14}H_9Cl_5$. In a 0.750 gram sample:

a. How many moles of $C_{14}H_9Cl_5$ are present?

b. How many grams of carbon are present?

c. What is the total number of atoms present?

d. What is the percent hydrogen in $C_{14}H_9Cl_5$?

8. How many grams of Na has the same number of atoms as 13.0 g N?

9. 1.450 moles of element Y weighs 0.30044 kg.

a. What is the molar mass of Y?

b. What element is this?

10. Write balanced chemical equations for the following reactions (you must include physical states)

a) When solid phosphorus is burned in oxygen, solid diphosphorus trioxide is produced

b) Solid barium carbonate and aqueous Ammonium chloride is produced **from** solutions of Barium chloride and ammonium carbonate

c) When Solid Iron (III) oxide is added to carbon monoxide gas, iron metal and carbon dioxide is produced

d) Phosphorus acid is produced from Diphosphorus trioxide solid being added to water.

e) Dinitrogen pentoxide + water → nitric acid

4. Balance the following :

$$Mg_2C_3 + H_2O \rightarrow Mg(OH)_2 + C_3H_4$$

REMEMBER TO DO THE STARRED PROBLEMS IN THE TEXTBOOK