CHEMISTRY 110 LECTURE

EXAM II Material

Part 1 ATOMIC STRUCTURE

I. The Atom - history:

A. Democritus (approx. 400 BC) - Mental concept of the atom: A repeated dividing of matter would eventually result in indivisible, invisible, minute particles called *atomos*

B. Dalton (approx. 1800 AD)- Experimentally based: Atoms are the building blocks of matter

C. Thompson (1897) Discovered the electron

D. Rutherford (1911) Developed the structure of the atom

E. Schrodenger (1923) Developed the orbital model of the atom/ electrons move in a wave motion

F. Chadwick (1932) Discovered the neutron

II. Subatomic Particles: -

III. ISOTOPES-Tools

A. Tools

1. Atomic number, $\mathbb{Z}$,, equals the number of protons

2. Mass number, $\mathbb{A}$, 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:

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. Give the isotopic symbol for an atom with 2 more neutrons and 1 more proton than P-30

C. IONS-Ions are atoms with a charge

IONS

ANIONS CATIONS

Ion-Problem

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?

IV. 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. Electron structure

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.

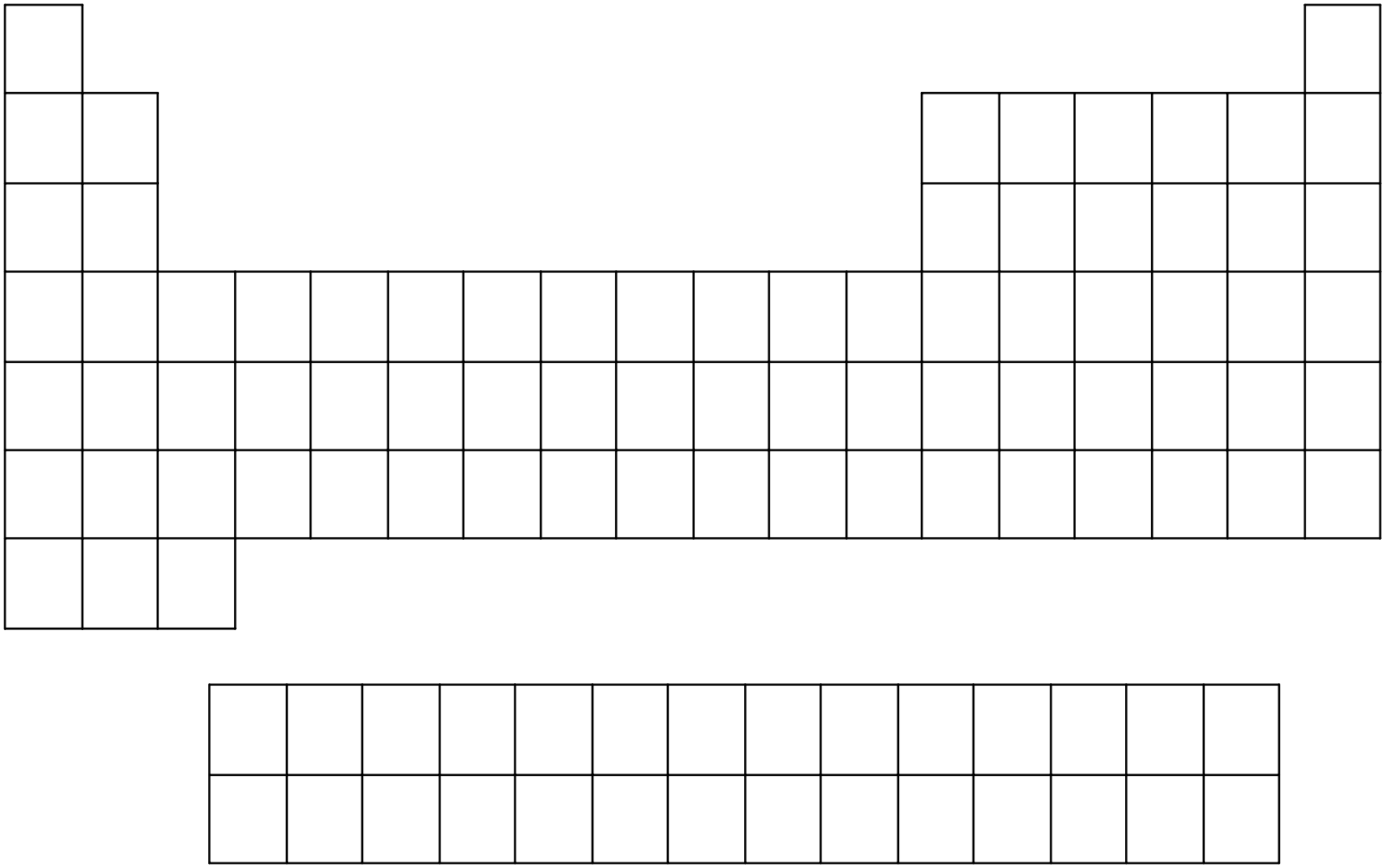
   b] Sublevel {s, p, and d} areas in the periodic table.

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

   (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.....



PERIODIC TABLE IA

Examples:

1. O

2. Si

3. Br

4. Na^{+}

5. S^{2-}

5. Ne

Mg

d. CORE NOTATION

Ca = [Ar]$4s^2$

↑ the "core" must be the previous inert gas

Examples:

Part 2 CHEMICAL BONDS

→ The attractive interaction between two atoms or ions

I. Types: -

1. **Ionic Bond**- Cations (+ charged) and Anions (- charged) 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. Valence electrons (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. Lewis Electron Dot Structures

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

IV. Octet Rule

Eight Valence electrons

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

Anions

Octet rule continued

Cations

V. Periodic Trends

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

A. Atomic Radius [Atom Size]

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

Summary

b. 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.

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

VI. Ionic Bonds - The attractive fore between a cation (+ ion) and anion (- ion).
Atoms lose or/ gain electrons to obtain an octet.

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

Part 3 NOMEMCLATURE

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:

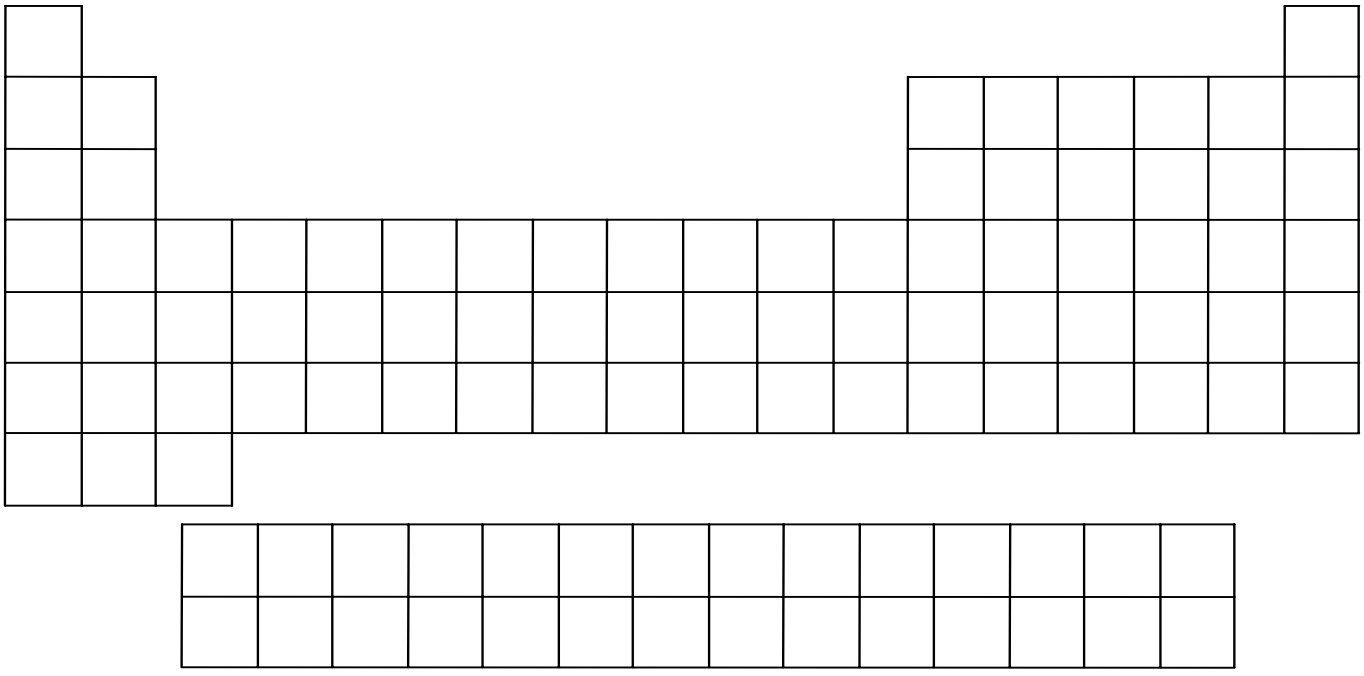
Metals form cations (+) ions:

Nonmetals form Anions (-) ions:

III. Cation and Anions

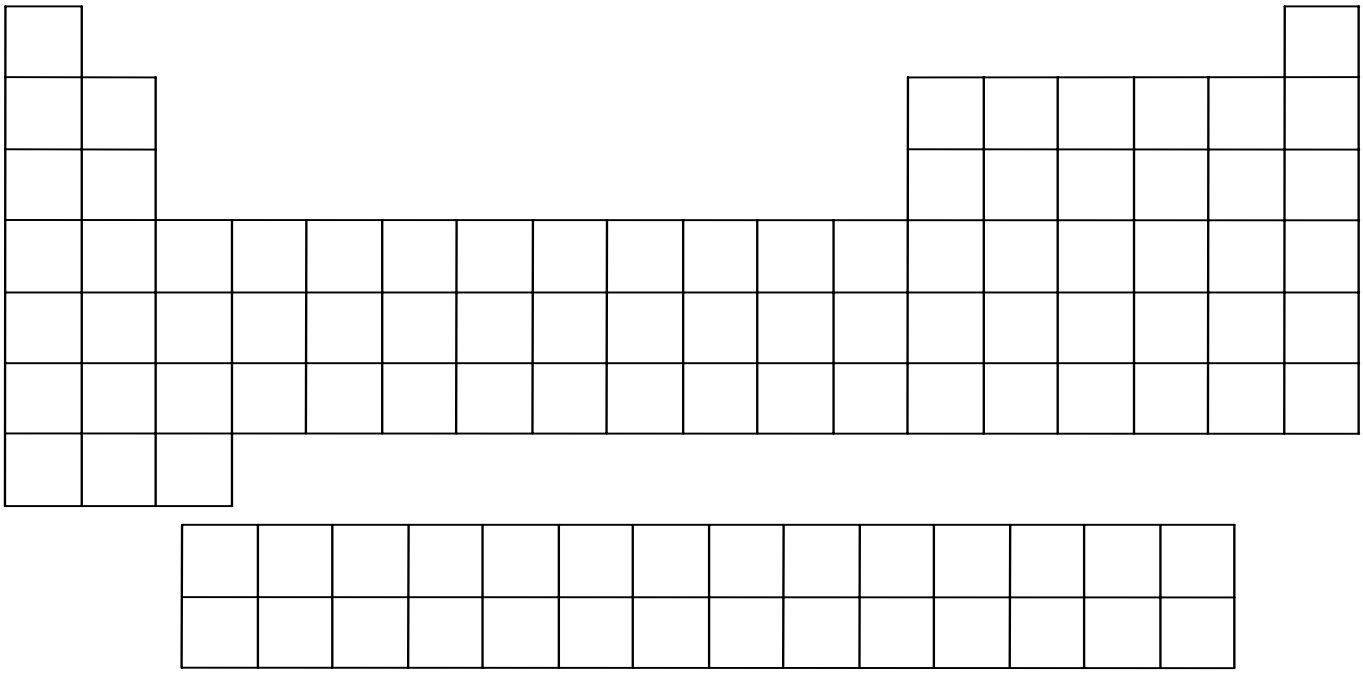
A. Metallic Cations - (+ charge)

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



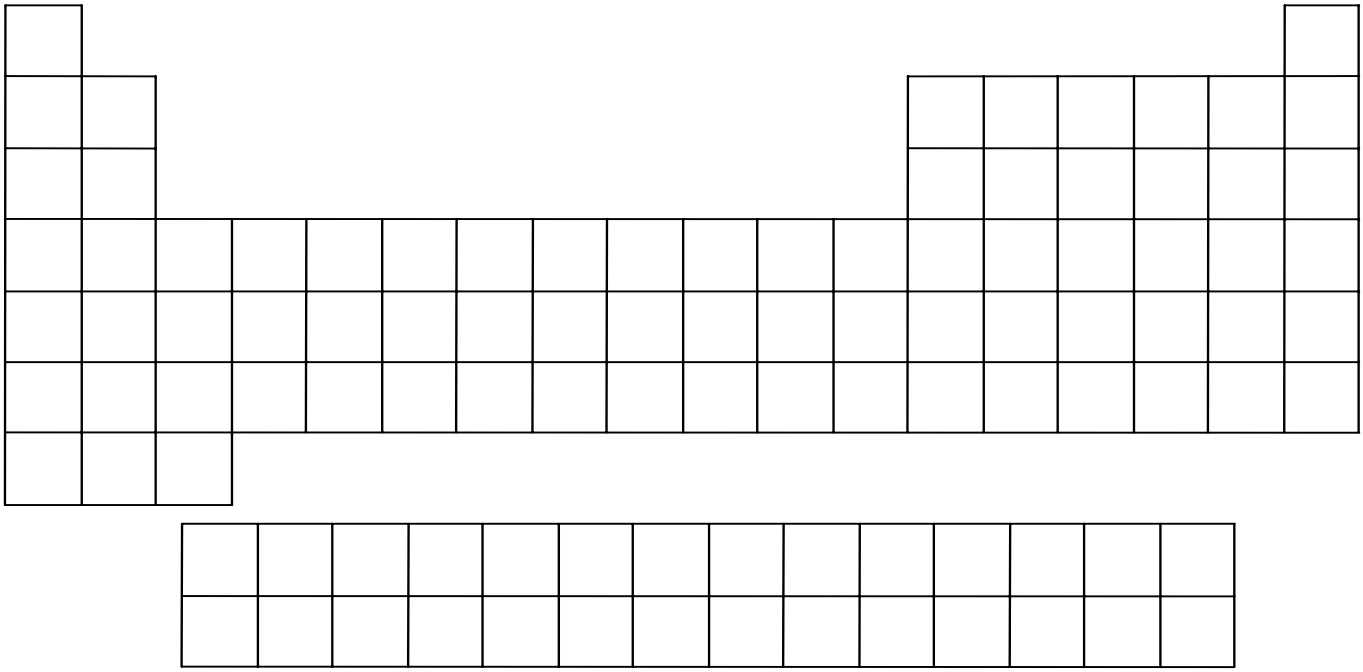
PERIODIC TABLE IA

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



PERIODIC TABLE IA

B. Nonmetal Anions (-) charge



PERIODIC TABLE IA

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/TERNARY ACIDS (contains "O")

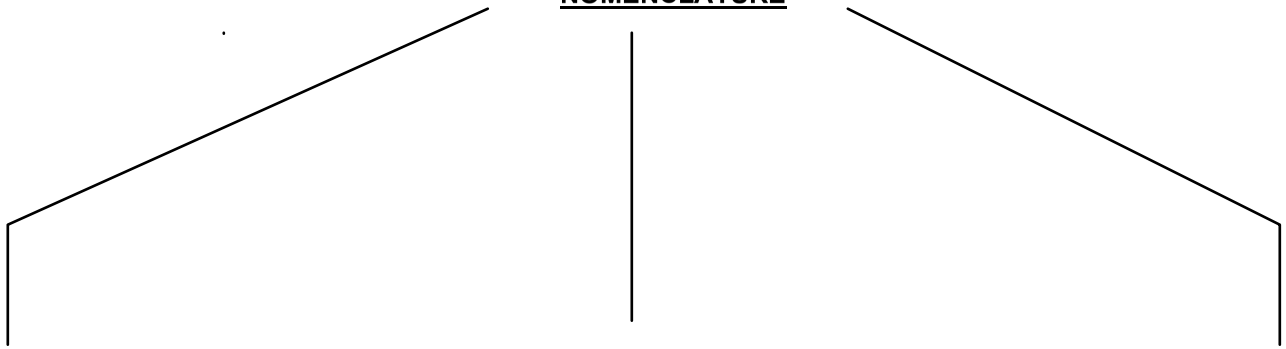
Naming Formula:

Ion name Bu<u>t</u> Change ite--> ous
ate--> ic
+ Acid

KEY: Recognize the ion part of the Acid

ACID	ION

NOMENCLATURE



Formula starts with an "H" + (aq)

(Water is excluded)

ACID

Metal-Nonmetal

IONIC COMPOUND

Nonmetal-Nonmetal

MOLECULAR

COMPOUND

oxyacid	binary acid	Fixed charged metal	Variable charged metal	When naming, use prefixes
contains oxygen	does not contain oxygen	Group IA,IIA, Al, Ga, Cd, Zn, or Ag --> name as is	All other metals -->place charge in ( ) as roman numerals/ know classical names	(mono, di tri...etc.)
EXAMPLES:				
HClO hypochlorous acid	HCl hydrochloric acid	KBr Potassium bromide	CuOH Copper (I) hydroxide	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

Part 4 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 x 10^{23} particles
Avogadro's number → memorize!!

Conversions

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

Problem: How many Cu atoms in 6.0 mol Cu?
Know: 1 mol Cu = 6.02 x 10^{23} atoms Cu

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 x 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	2 mole N
5 atoms O	5 moles O
= 1 molecule N_2O_5	= 1 mole of N_2O_5

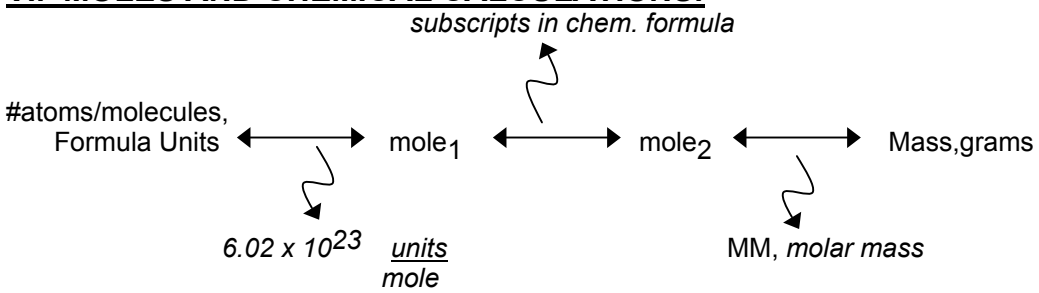
Ratios:

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:



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 **<u>total</u>** number of atoms in 0.20 mole 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.

8.

9.

10.

VII. PERCENTAGE COMPOSITION

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

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.

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:

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

Practice EXAM IIA

100 POINTS

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	Co_2O_3
$Ni(NO_2)2$	$Bi(NO_3)_3$
Ba_3N_2	$HClO_3(aq)$
$Ca(OH)_2$	N_2O_5
$Sr(HSO_3)_2$	$Hg(HCO_3)_2$
$H_2CO_3(aq)$	PbO_2
SO_3	
HF	$HBrO_2(aq)$
$HC_2H_3O_2(aq)$	Au_3PO_4
N_2O_3	$Cu(Cl0)_2$
HCN(aq)	$Al(OH)_3$
KH	

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 <u>total</u> 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. For the following questions , identify the element whose atoms fit the following descriptions **(Use the periodic table)**

a. ____ Which has d electrons?

a) hydrogen b) Copper c) nitrogen

b. ____ A metalloid in period 5

c. ____ The element in period 4, group IIIA

d. ____ The element with a total of 3 electrons in the 2nd main energy shell

e. ____ The element with 5 completely filled orbitals

f. ____ The smallest alkali metal

g. ____ The largest period 4 transition metal

11. a. Write the electron configuration of a potassium ion

   b. Write the electron configuration of tin

12. Write the electron dot structure for the following compounds

$H_2SO_3(aq)$ NO_3^-

REMEMBER TO DO THE STARRED PROBLEMS IN THE TEXTBOOK