Chemistry 101

General Chemistry I

Course Syllabus

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Required Materials

- Textbook: General Chemistry, 9th Ed. Ebbing and Gammon, © 2009.
- Laboratory Manual: Lab Manual for Chemistry 101, Dept. of Chemistry, LAVC, © 2008. Download from website.
- Laboratory Safety Manual from ACS website.
- Scientific calculator (must be capable of scientific notation and logarithms)
- Chemical Safety goggles

Course Grading

- 800 points
  - 300 points for exams
  - 100 points for the safety quiz
  - 200 points for the final
  - 200 points for lab
- No Curve, Straight scale (90%, 80%, etc.)
  - Grade is based solely on the percentage of points, not the number of points.
- Start worrying about your grade today!
- Each day study as if you have to take the final tomorrow.

Course Grading (cont'd)

- You have the chance to make up a bad exam with the final exam. If the percentage on the final exam is higher than the lowest of your exams it will replace it. Example:
- Exam 1: 87% Exam 2: 93%
- Exam 3: 67% Exam 4: 90%
- Scenario 1: Final exam: 88% – The 88% will replace the 67%.
- Scenario 2: Final exam 57% – The 57% will **not** replace the 67%.

Final Exam

- Final exam will be on …

14 December 2009 at 9:30 a.m.

What are you expected to know?

- All of the information in Chemistry 68…
  - Significant figures
  - SI Units
  - Metric prefixes
  - Derived units (e.g. density)
  - Key Terms
  - Subatomic particles and the nuclear atom

You are expected to be skilled in an routinely use dimensional analysis!

Homework

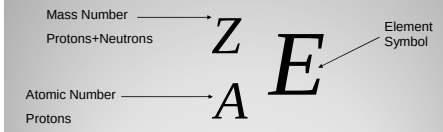
- Other than labs there is no homework to be turned in.
- In your syllabus there are several Suggested Problems (highly suggested)
- These are for your benefit. Do them! Ask questions about them!

Atoms

All about atoms

Atomic Symbols

- Information about the atom is summarized in the atomic symbol.



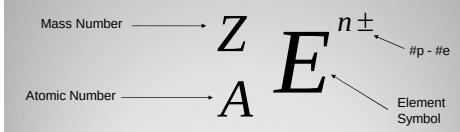
Isotopes

- Isotopes are elements with the same atomic number but different mass numbers.
- An element can have one or more isotopes.
- Elements with the same mass number but different atomic numbers are called ***isobars***.

Isotopes

- Each isotope occurs in nature with a certain abundance (can be 0%).
- The abundances are not the same throughout the universe.
  - For example Iridium is more abundant in asteroids than in the earth.
  - The isotopic ratios are also different from place to place

- Ions are atoms that have lost or gained electrons
- We can represent ions this way…



Ions

- The mass number is not the same as the mass.
- Mass numbers are always integers. Masses are not (look at the periodic table).
- Why is this?

Masses of elements

- The numbers in the table are ***relative average atomic masses***.
- Average over all of the isotopes, which have different masses.
- Relative to what?
  - Carbon-12 which has a mass of exactly 12 amu, by definition.

Relative average atomic masses

- You need the isotopic composition.
- Convert the percent abundances to decimal.
- Multiply each abundance by its respective mass.
- Add all of the above products.

How to calculate?

Example

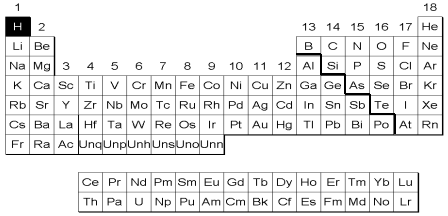
- Indium has two naturally occurring isotopes.

Isotope	%	Mass (amu)
$^{113}_{49}In$	4.3	112.9041
$^{115}_{49}In$	95.7	114.9057

- Convert percentages to decimal.
  - 4.3% → 0.043
  - 95.7% → 0.957
- Multiply each decimal abundance by it mass.

$$(0.043)(112.9041\text{ amu})+(0.957)(114.9057\text{ amu})$$
$$=4.85487\text{ amu}+109.96475\text{ amu}$$
$$=115\text{ amu}$$

Solve

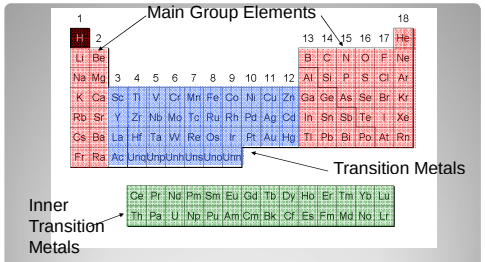


What are the features of the periodic table?

The Periodic Table

- Vertical columns are called groups or families
  - Special groups
    - Group IA ⇒ Alkali metals
    - Group IIA ⇒ Alkaline Earth metals
    - Group VIIA ⇒ Halogens
    - Group VIIIA ⇒ Noble Gases
- Horizontal rows are called periods
- "Stair-step" line separates the metals on the left from the non-metals on the right. It's bordered by the metalloids or the semi-metals.

Features



More Features

- Ions from metals always have positive charges (lose electrons).
  - Charge will equal group number for most main group metals.
- Ions from non-metals always have negative charges (gain electrons).
  - Charge will equal group number minus 8.

Main Group Elements

The Mole

Counting atoms and molecules

- How many atoms are in a cubic meter of lead?
- To answer this we need to know the mass of the m^3 of lead. The mass will tell us the number of atoms indirectly.
- The mass tells us the number of ***moles***.

Counting atoms

What is a mole?

- It's a counting number just like a pair or a dozen or a ream.
- It's defined as the number of atoms in exactly 12 g of Carbon-12.
- It's a very, very large number.
  - $6.02214199 \times 10^{23}$

Comparisons…

- 1 mole of dollars
  - You could spend $1,000,000 per second for the rest of your life and only spend 0.4% of the money
- 1 mole of sand
  - 1 mole of sand would cover Los Angeles to a depth of 600 m.

A Mole is a convenient unit

- 1 atom of oxygen is 15.9994 amu
- 1 mole of oxygen atoms is 15.9994 g.
- This relationship is true for all atoms and molecules.

Compounds and Their Formulas

Molecular Compounds

- Molecular compound are made of discreet particles called molecules.
- They are held together, internally, by covalent bonds, or shared electrons.
- The chemical formula gives the number of each type of atom in the molecule.
  - $C_2H_6O \rightarrow$ 2 Carbons, 6 Hydrogens and 1 Oxygen

Ionic Compounds

- There are no discreet molecules in ionic compounds.
- They're composed of ions
  - Positive ions are called cations
  - Negative ions are called anions
- The chemical formula gives the ratio of cations to anions.
  - $FeCl_2 \rightarrow$ 1 Iron cation to 2 Chloride anions

Masses of Compounds

- Formula mass – Applies to all types of compounds
- Molecular mass – Applies only to molecules
- Either one is the sum of the masses of all the atoms in the compounds formula.

Example

- What is the molecular mass of Hydrogen Peroxide (H_2O_2)?
  - 2 Hydrogen atoms and 2 Oxygen atoms

$$2(1.008\ \text{amu}) + 2(15.9994\ \text{amu}) = 34.01468\ \text{amu}$$
$$= 34.0147\ \text{amu}$$

Molar mass

- The mass of one mole of a substance.
- Numerically the same as the formula mass but the units are grams instead of amu.
- Can be used as a conversion factor between mass and moles of a substance.
  - $\frac{1\ \text{mol}\ H_2O}{18.0153\ \text{g}\ H_2O}$ or $\frac{18.0153\ \text{g}\ H_2O}{1\ \text{mol}\ H_2O}$

Elements as Molecules

- Several elements exist as molecules in their standard states.

(Memorize!!!)

- H_2
- N_2
- O_2
- F_2
- Cl_2
- Br_2
- I_2
- P_4
- S_8

Determining Chemical Formulas

- In order to determine the chemical formula of a compound we need to know what the percent composition of the compound is.
- The percent composition is the percentage of the mass of the compound that each element comprises.
  - Example: Water is approximately 11.2% Hydrogen and 88.8% oxygen.

Determining Chemical Formulas

- The formula that is determined from the percent composition, or any other data, is the empirical formula.
- An empirical formula is a formula that gives the lowest whole number ratio of atoms in the compound.
  - Formulas for ionic compounds are always empirical formulas
  - Formulas for molecular compounds may or may not be empirical formulas.

Empirical Formulas

- Examples
  - Water: Molecular formula: H_2O,
    Empirical formula: H_2O
  - Hydrogen Peroxide: Molecular Formula: H_2O_2
    Empirical Formula: HO
  - Glucose: Molecular Formula: $C_6H_{12}O_6$,
    Empirical Formula: CH_2O

Determining Chemical Formulas

- From the percentages we find the mass of each element present (assume 100.00 g of compound.
- Convert each of the masses into moles.
- Divide each of the moles by the smallest number.
  - The numbers we get are the subscripts in the formula.
  - If the number is not within ±0.1 of an integer we must multiply by an integer to get close to an integer.

Example: Nicotine

- Nicotine is 74.03% C, 8.70% H, and 17.27% N by mass. What is the empirical formula of Nicotine?
  - Convert each of the percentages to masses:
    - 74.03 g of C
    - 8.70 g of H
    - 17.27 g of N

Example: Nicotine

- Convert each of the percentages to masses:
  - 74.03 g of C
  - 8.70 g of H
  - 17.27 g of N
- Convert each mass to moles

Example: Nicotine

- Convert masses to moles:

$$74.03\text{ g C}\times\frac{1\text{ mol C}}{12.0107\text{ g C}}=6.16\underline{3}670726\ldots\text{ mol C}$$

$$8.70\text{ g H}\times\frac{1\text{ mol H}}{1.00794\text{ g H}}=8.6\underline{3}138738\ldots\text{ mol H}$$

$$17.27\text{ g N}\times\frac{1\text{ mol N}}{14.0067\text{ g N}}=1.233\text{ mol N}$$

Example: Nicotine

- Divide each by the smallest number of moles:

$$6.164\text{ mol C}/1.233=4.996\approx 5$$

$$8.631\text{ mol H}/1.233=7.000=7$$

$$1.233\text{ mol N}/1.233=1.000=1$$

The empirical formula is C_5H_7N.

Finding Molecular Formulas

- We need to divide the molecular mass by the empirical mass. This ***will*** give us an integer.
- The subscripts are multiplied by this integer to get the molecular formula.

- Hydrocarbons and derivatives of hydrocarbons can be analyzed with this method.
- Burn the substance in O_2 and measure the mass of CO_2, H_2O and other things that come out.
- O in the compound is determined by ***difference***.

Combustion Analysis

- Aniline is a starting compound for the productions of various dyes. 0.922 g of aniline was burned and produced 2.565 g of CO_2, 0.525 g of water and 0.163 g of N_2 gas. What is the empirical formula for aniline?

Example

- Convert g of CO_2, H_2O and N_2 to g of C, H, and N respectively.

$$2.565\text{ g } CO_2 \times \frac{1\text{ mol } CO_2}{44.01\text{ g } CO_2} \times \frac{1\text{ mol C}}{1\text{ mol } CO_2} = 0.05828\text{ mol C} \times \frac{12.0107\text{ g C}}{1\text{ mol C}} = 0.7000\text{ g C}$$

$$0.525\text{ g } H_2O \times \frac{1\text{ mol } H_2O}{18.015\text{ g } H_2O} \times \frac{2\text{ mol H}}{1\text{ mol } H_2O}$$

$$= 0.05828\text{ mol H} \times \frac{1.00794\text{ g H}}{1\text{ mol H}} = 0.05874\text{ g H}$$

$$0.163\text{ g } N_2 = 0.163\text{ g N} \times \frac{1\text{ mol N}}{14.0067\text{ g N}} = 0.01166\text{ mol N}$$

Solution

- Subtract g of C, H and N from total mass to find mass of O.
- There is no O in this compound.

$$0.922 - 0.7000 - 0.05874 - 0.163 = 2.0 \times 10^{-4} \approx 0$$

Solution

- Now we divide the moles of atoms by the smallest, just like before.

$$\frac{0.05828\text{ mol C}}{0.01166\text{ mol}} = 4.998\text{ C} = 5\text{ C} \qquad \frac{0.05828\text{ mol H}}{0.01166\text{ mol}} = 4.998\text{ H} = 5\text{ H}$$

$$\frac{0.01166\text{ mol N}}{0.01166\text{ mol}} = 1.000\text{ N}$$

- Therefore the empirical formula is C_5H_5N.

Solution