Basic Chemistry Concepts

Lecture Notes on Laws of Chemical Combination and Atomic and Molecular Masses

Laws of Chemical Combination

1. Law of Conservation of Mass

- **Definition**: States that mass is neither created nor destroyed in a chemical reaction. The mass of the reactants equals the mass of the products.
- **Example**: When hydrogen and oxygen react to form water, the total mass of hydrogen and oxygen before the reaction is equal to the mass of water after the reaction.

2. Law of Definite Proportions

- **Definition**: States that a given chemical compound always contains its component elements in a fixed ratio by mass, regardless of the source or method of preparation.
- **Example**: Water (H_2O) always contains hydrogen and oxygen in a mass ratio of 1:8.

3. Law of Multiple Proportions

- **Definition**: States that if two elements combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element are in the ratio of small whole numbers.
- **Example**: Carbon and oxygen form CO and CO_2. The mass of oxygen that combines with 1 gram of carbon in CO_2 is twice that in CO, giving a simple ratio of 1:2.

4. Law of Reciprocal Proportions

- **Definition**: States that if two elements, A and B, combine separately with a fixed mass of a third element, C, the ratio of the masses in which they do so will be the same or a simple multiple if A and B combine directly with each other.
- **Example**: If 8 grams of oxygen combine with 1 gram of hydrogen to form water, and 8 grams of oxygen combine with 3 grams of carbon to form carbon dioxide, then hydrogen and carbon combine in a simple ratio of 1:3.

5. Gay-Lussac's Law of Gaseous Volumes

- **Definition**: States that when gases react together at constant temperature and pressure, the volumes of the reacting gases and the volumes of the products (if gaseous) are in simple whole number ratios.
- **Example**: When hydrogen and oxygen react to form water vapor, 2 volumes of hydrogen react with 1 volume of oxygen to produce 2 volumes of water vapor.

6. Avogadro's Law

- **Definition**: States that equal volumes of gases, at the same temperature and pressure, contain an equal number of molecules.
- **Implication**: This law allows the determination of relative molecular masses of gases from the volumes of gases that react.

Atomic and Molecular Masses

1. Atomic Mass Unit (AMU)

- **Definition**: A unit of mass used to express atomic and molecular weights, defined as one twelfth of the mass of a carbon-12 atom.
- **Symbol**: u or amu

2. Calculation of Atomic Mass

- **Definition**: The atomic mass of an element is the weighted average mass of the atoms in a naturally occurring sample of the element.
- **Calculation**: Sum of the masses of protons, neutrons, and electrons in an atom. However, the mass of electrons is negligible.
- **Example**: The atomic mass of carbon is approximately 12.01 amu.

3. Calculation of Molecular Mass

- **Definition**: The molecular mass (molecular weight) of a molecule is the sum of the atomic masses of all the atoms in the molecule.
- **Calculation**: Add the atomic masses of all the atoms in the molecule.
- **Example**: The molecular mass of water (H_2O) is calculated as (2 × 1.008) + 16.00 = 18.016 amu.

4. Average Atomic Mass

- **Definition**: The average mass of atoms of an element, calculated using the relative abundance of isotopes in a naturally occurring element.
- **Calculation**: (Fractional abundance of isotope 1 × mass of isotope 1) + (Fractional abundance of isotope 2 × mass of isotope 2) + ...
- **Example**: For chlorine, with isotopes Cl-35 and Cl-37:

$$\text{Average atomic mass} = (0.75 \times 34.97) + (0.25 \times 36.97) = 35.47\ \text{amu}$$

These notes cover the fundamental concepts of the laws of chemical combination and atomic and molecular masses, providing essential knowledge for understanding chemical reactions and the properties of elements and compounds.