Chemistry IH - Study Guide

Mr. Kohl - 2nd Hour

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Chapter 1

Introduction

1.1 Types of Chemistry

Inorganic Chemistry relates to the chemistry of molecules that do not contain a carbon atom

Organic Chemistry the chemistry of molecules which do contain a carbon atom. These molecules may also contain a hydrogen atom, and always contain carbon bonds.

Biochemistry the chemistry of proteins and large bio-molecules

Analytical evolving methods to identify and measure things (relating to chemistry)

Theoretical considered the *opposite* of analytical chemistry. Theoretical chemistry is focused on trying to preduct things about chemistry using calculus.

Physical uses the theory developed from the above discipline to compare it against what *actually happened*

1.2 Types of Matter

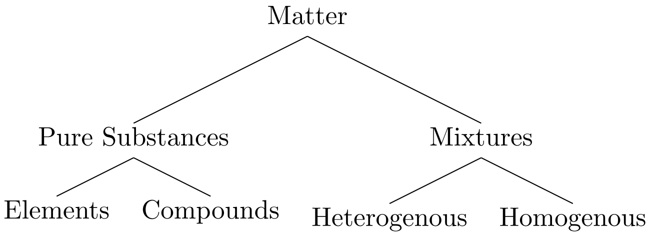


Figure 1.1: The types of matter.

For each type of matter, a definition and several real-world examples are found below.

Elements matter that cannot be broken down into more pure substances *i.e., they have but only one unique type of atom*

- Carbon
- Hydrogen
- Oxygen

Compounds pure substances made ov several elements of which are *chemically* joined together

- H_2O
- CO_2

Mixtures impure combinations of two or more substances and are *not* chemically joined

Homogenous uniform content throughout

- Alloys of metal
- Filtered air

Heterogenous non-uniform content throughout

- Non-pasturized apple juice

1.3 Separation of a Mixture

Filtration used on a heterogenous mixture by means of a porris barrier. Exploits the physical size difference *ex., using a coffee filter to make coffee.*

Distillation used on homogenous mixtures and takes advantange of temperature differences. *ex., alcohol boils at temperature A, water boils at temperature B. A temperature between the two is picked to boil off one compound, and then condense it later.*

Chromatography used on both types of mixtures to determine properties of compounds to seperate. In practice, a mixture is placed on the end of a piece of paper, and water draws out all the compounds contained within it.

1.4 General Properties of Substances

Intensive properties of substances which are not related to the amount of that substance which is present

- Boiling point
- Specific heat
- Density

Extensive properties of substances which are functions of the amount of that substance which is present

- Mass
- Volume
- Quantity of atoms

1.4.1 Chemical vs. Physical

Chemical observed when something changes into a new substance

- Bond
- Toxicity
- Reactivity
- Flamability

Physical observed without changing chemical properties

- Color
- Density
- Temperature

1.5 SI Units

The SI-system is the unit of measurements most widely used by scientists. It defines a prefix-base system which is useful in representing both large and small numbers.

1.5.1 Base Units

Length meter m

Mass kilogram kg

Time second sec

Teperature kelvin k

Amount of Substance mole mol

1.5.2 Derived Units

Name	Symbol
Area	m^2
Volume	m^2
Density	kg/m^3
Molar Mass	kg/mol
Concentration	mol/L
Molar Volume	L/mol
Velocity	m/s
Force	m/s^2

1.5.3 Prefix System

	Symbol	Modifier
mega	M	10e6
kilo	K	10e3
centi	c	10e−2
milli	m	10e−3
micro	μ	10e−6
nano	n	10e−9

1.5.4 Significant Figures

Significant Figures define a set of rules for maintaing percision while doing math. This system ensures that the answer to a problemset accurately reflects the degree of percision of all sub-compounds.

1. Zeros between non-zero digits are considered significant
2. Zeros in front of a number are not significant
3. Zeros to the end of a number and to the right of a decimal place are considered significant
4. Zeros to the right of non-zeros *may* be considered significant. 1000 has one digit of significance, while 1000 has four. Similarly, 1 00e3 has three digits of significance.

1.5.5 Rules for Arithmetic

When using digits of significance in arithmetic problems, it is important to maintain a correct number of significant digits in the answer. When adding and subtracting numbers, the result has the minimum of the digits of significance between the two products. The same rules apply to both multiplication and division.

Chapter 2

Atoms and Elements

2.1 History

Democritus thought that all things are made of indivisible bits

Aristotle disagreed with the idea of the atom and muted discussion

2.2 Basic Laws

The Law of Conservation of Mass matter is neither created nor destroyed during ordinary chemical and physical changes.

The Law of Definite Proportions chemical compounds have the same elements in the same proportions by mass regardless of the size of the sample

The Law of Multiple Proportions when chemical elements combine they will do so in small, whole-number rations. If two elements can form multiple compounds, the ratio of the second element to a fixed mass of the 1st element will always be a small, fixed number.

2.3 Dalton's Atomic Theory

1. All mater is composed of atoms
2. Atoms of an element are identical, atoms of different elements are dissimilar in proportions
3. Atoms cannot be subdivided, created, or changed
4. Atoms of different elements combine in simple whole number ratios to form chemical compounds
5. In chemical reactions, atoms are combined, seperated or rearranged

2.4 Modern Atomic Theory

2.4.1 Universal Truths

- Conservation of mass is obeyed during chemical reactions
- All mater is composed soley of atoms
- Atoms of any one element differ in proportions from others

2.4.2 Variations on Dalton's Theory

- Atoms of a type can have different masses (isotopes)
- Atoms have an internal structure composed of protons, neutrons, and electrons
- Atoms can be created and destroyed, but only during nuclear reactions

2.5 Modern Structure of an Atom

An atom is the smallest unit of a element which still retains the properties of that element.

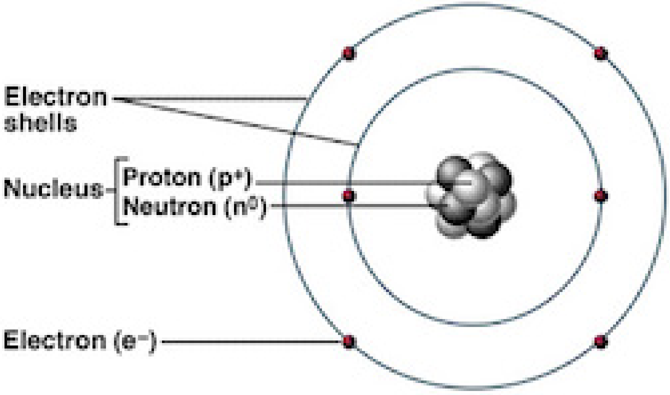


Figure 2.1: Modern-day structure of an atom (basic)

There are three components that make up the important parts of an atom. They are described below:

Nucleus The center of the atom. Very small, and very dense. Contains protons and neutrons.

Protons positively charged sub-atomic particles

Neutrons neutrally charged sub-atomic particles

Electrons Small, negatively-charged particles. They orbit the nucleus in *electron clouds*. That is to say that we do not know the definite position of an electron, but we can reason about where it is likely to be.

2.6 Experiments

2.6.1 Cathode Ray Tube

Experiment preformed by J.J. Thompson. Contributed first to the discovery of a sub-atomic particle (the electron). Showed that a cathode-ray had negatively charged particles because it responded to magnetic fields.

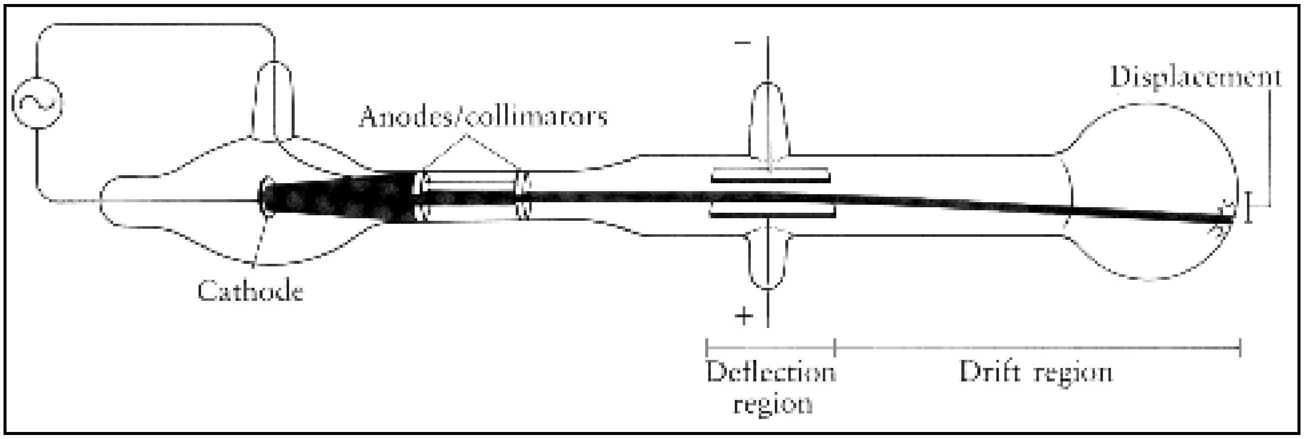

Anodes/collimators
Cathode
−
+
Displacement
Deflection region
Drift region

Figure 2.2: Diagram of the cathode-ray tube expirement.

2.6.2 Milikan Oil Drop

Determined the more specific charge of the electron. Showed that small particles were able to be suspended when they picked up negative charge. Using the charge of the plates involved, the charge was able to be determined.

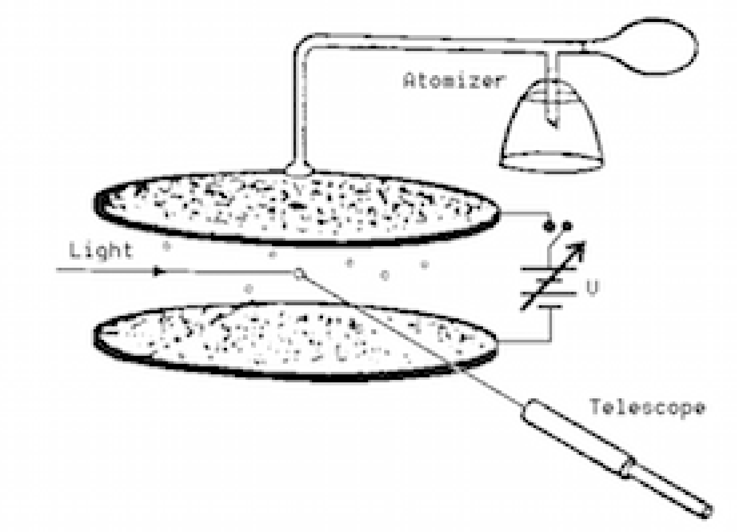

Atomizer
Light
U
Telescope

Figure 2.3: Diagram of the oil-drop experiment.

2.6.3 Rutherford's Gold Foil

Experiment preformed by Rutherford and Geiger. Together, the two of them shot *alpha* particles at a ultra-thin sheet of gold foil. These particles exhibited one of three behaviors, each indicating a different outcome.

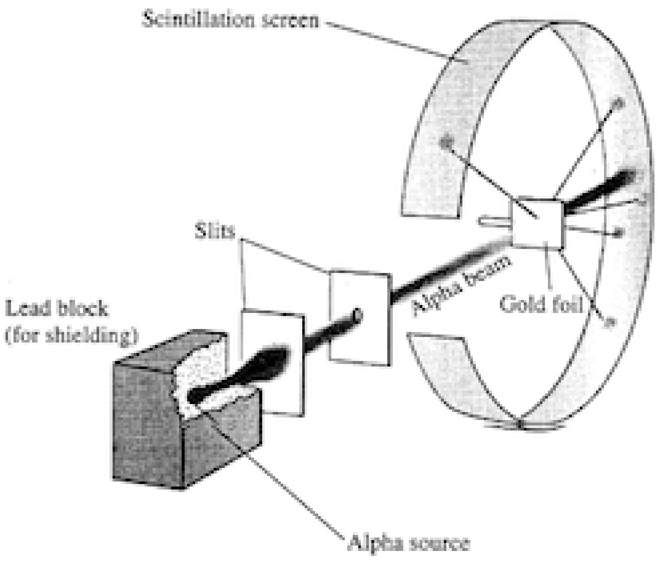


Figure 2.4: Diagram of Rutherford's Gold Foil experiment.

Through the foil the alpha particles went through empty space within an atom, and thus traveled uninterrupted through the atom in a straight path

Deflected through the foil the alpha particles were attracted to something while traveling through an empty space within the atom. As such, their path was changed slightly, and they angled as they travelled.

Bounced off the foil the alpha particles collided with the positively charged nucleus to deflect off of the foil, back at the particle gun.

2.7 Elements

Elements are defined by the number of protons in the nucleus of one of their atoms. No two elements will ever have the same number of protons in their nucleus under normal circumstances. *As such, the number of protons in an element can be considered as its unique key.*

The atomic number of an element has the same value as the number of protons in its nucleus.

2.7.1 Isotopes

Isotopes have the same number of protons, but a different number of neutrons. Isotopes are represented as the following $^{35}_{17}\mathrm{Cl}$. This is Chlorine-35. The superscript represents the atomic mass of an isotope, and the subscript represents the atomic number.

Since the atomic mass is the sum of the number of protons and the number of neutrons in an atom, the number of neutrons in an isotope can easily be determined.

2.8 Nomenclature

2.8.1 Common Nomenclature

Molecular Compounds compounds which contain only non-metal components

Binary Compounds compounds of two different elements. In non-ionic binary compounds, simply change the last component to have an *-ide* suffix, for example $CaCl_2$ - *Calcium Chloride.* Naming occurs in the order *mono-, di-, tri-, tetra-, penta-, hexa-, hepta-, octa-, nona-, deca-.*

Binary Ionic Compounds binary-compounds (see above) which contain both a metal and a non-metal.

1. Positive charge first
2. The positive element undergoes no change in the name of its element
3. Negative charge gets the *-ide* treatment

2.8.2 Stock-system Nomenclature

When dealing with ions, it is frequently the case where there are more than one charge state available to work with. The stock-system nomenclature makes it obvious which one we are dealing with.

Simply put, you take the charge of an ion, and place it in parenthesis in roman-numerals to the right of the chemical symbol. For example, Fe^{3+} becomes Fe(III).

2.9 Types of Ions

Ions take shape in several different types, but come in only two charges. These charges are named *cation* and *anion.* A cation is a positively charged ion, whereas an anion is a negatively charged ion. *(Remember: a-negative-ion... an-ion.)*

2.9.1 Monatomic Ions

Monatomic ions are formed of one ore more of the same kind of element. These ions gain and loose electrons to obtain a full valence-level of electrons for wahtever level they might be at. Examples include H^+ and O^{2-}.

2.9.2 Polyatomic Ions

Polyatomic ions are charged groups of covalently bonded atoms. They carry charge as a group, such that they behave like single ions, even though they are composed of more than one atom.

All relevant polyatomic ions are listed below, grouped by charge.

- Cation^{1+}

  Ammonium ${NH_4}^+$
  Hydronium ${H_{30}}^+$

- Anion^{1-}

  Cyanide CN^-
  Bicarbonate ${HCO_3}^-$
  Hydroxide OH^-
  Nitrate ${NO_3}^-$
  Nitrite ${NO_2}^-$
  Permanganate ${MnO_4}^-$
  Acetate CH_3COO^- or ${C_2H_3O_2}^-$
  Dihydrogen Phospate ${H_2PO_4}^-$

- Anion^{2-}

  Carbonate CO_3^{2-}
  Sulfate SO_4^{2-}
  Sulfite SO_3^{2-}

- Anion^{3-}

  Phosphate PO_4^{3-}

Chapter 3

Atomic Structure and Periodic Trends

3.1 Electron-Configuration Notation

A system of numbers and letters is used to designate electron configuration. For example: $1\,\mathrm{s}^2 2\,\mathrm{s}^2 2\,\mathrm{p}^6 3\,\mathrm{s}^2 3\,\mathrm{p}^2$.

1. the level number is used
2. the level designation for the shape of the orbital
3. a superscript is used to represent the number of electrons in that specific orbital.

3.1.1 Noble-gas Notation

To simplify writing out all of these numbers, you can include the noble-gas most-closely following behind of the element you are trying to describe, and then only describe the differing electrons using the notation above. For example: $\mathrm{Ne} 3\,\mathrm{s}^2 3\,\mathrm{p}^2$.

Highest occupied level the electron containing main energey level with the highest quantum number

Inner-shell electrons the electrons that are contained in levels with a lower quantum state

Noble-gas configuration an outer main-energey level that is fully occupied, usually by eight electrons.

3.2 Periodic Table and Trends

3.2.1 The Periodic Table

Mendeleeve's Table

In 1860, the first Internation Conference of Chemists was held and at that time, the first periodic table was discussed. In 1869, 9 years later, that table was published in a textbook for college students.

Mendeleev was the first to generate a periodic table. He used atomic mass as the basis for his table, arranging elements by increasing mass. Some elements, however, were out of place when compared against the periodic tables of today. He was able to, despite these faults, see periodic trends, which allowed him to predict with a high-degree of accuracy, the qualities of yet undiscovered elements.

Improvements

Moseley observed that elements were better fit as ordered by their nuclear charge, not atomic mass. The nuclear charge is due to the number of protons, which led to the usage of atomic number for description.

Periodic Law states that the physical and chemical properties of the elements are periodic functions of their atomic numbers.

3.2.2 Modern-day

Today, the *Periodic Table* is just an arrangement of elements in the order of their atomic numbers, such that elements of similar properties fall in the samily column or group. However, there are but a few differences from Mendeleeve's original table:

1. There are more elements now as compared to the first introduction of the periodic table
2. The discovery of Noble Gases and the synthesis of the Lanthanides and Actinides
3. The general arangement has changed over time as well

3.3 Periodic Trends

3.3.1 Atomic Radii

Atomic radius is defined as one half of the distance between the nuclei of identical atoms that are bonded together. Atomic radii decreases as we move across a period (increase in positive charge) and increases as we descend a family (increasing the main energy level).

3.3.2 Ionization Energy

Ionization energy is the energy required to remove one electron from a neutral atom of an element. It is defined by the following chemical equation: $\mathrm{A} + \mathrm{energy} \longrightarrow \mathrm{A}^{+} + \mathrm{e}^{-}$. Ionization energy increases as you move right across a period, and decreases as you move down a family.

Multiple Ionization Energies

There are multiple ionization energies, each which involves removing another electron from the atom. These are referred to as IE_2, IE_3, and so on.

3.3.3 Electron Affinity

Electron affinity is the amount of energy change that results when an electron is acquired by a neutral atom.

3.3.4 Ionic Radii

This trend represents the radius of an ion of an element. Cations have ionic radii that are smaller than the neutral atom, where as anions have ionic radii that are typically larger than the neutral atom. This happens because of the addition and removal of electrons in the outer shell.

3.3.5 Electronegativity

Electronegativity is the measure of the ability of an atom in a chemical compound to attract electrons.

Chapter 4

Chemical Bonding

4.1 Introduction

A *chemical bond* is a mutual attraction between the nuclei and valence electrons of different atoms that binds the atoms together. Atoms form chemical bonds to gain a greater stability.

In the formation of a bond, atoms share, or give up, their valenece electrons. This sharing results in a greater stability as a whole for the system. The way in which the electrons are shared is determinant of the type of bond that is formed.

4.2 Ionic Bonds

Ionic bonding occurs from the electrical attraction between large numbers of positive, and negatively charged ions. The difference in electronegativity determines what type of bond it is.

Ionic difference in electronegativity between 1.7 and 3.3.

Molecular (Polar) difference in electronegativity between 0.3 and 1.7.

Molecular (Nonpolar) difference in electronegativity between 0 and 0.3.

4.3 Covalent Bonds

Covalent bonds represents the sharing of electron pairs between two atoms.

4.3.1 Non-polar covalent bonds

These bonds are covalent bonds in which the bonding electrons are shared equally between the bonded atoms, creating a balanced distribution of charge. They are 0 to 5% ionic character and have an electronegativity difference of 0 to 0.3.

4.3.2 Polar covalent bonds

Bonds that are 5 to 50% percent ionic in character have an uneven distribution of charge, such that the bonded atoms have an unequal attraction for the shared electrons.

4.4 Terminology

Molecule a neutral group of atoms that are held together by covalent bonds

Molecular compound a chemical compound where the simplest unit is molecules

Chemical Formula indicates the relative number of atoms of each kind in a chemical compound using atomic symbols and numerical subscripts

4.5 Motivation

Why form a bond? To answer this question, we must first understand the motivation of atoms. All atoms seek to become the most chemically stable that they can be, and if they can lose, gain, or share a valence electron in order to do so, they will.

However, bonded atoms behave not as one might initially assume. Two atoms which are far away initially have no force of attraction, however, as they get closer, this force increases and increases, until they are *too* close and thus begin to repel each other. As the atoms back off each other, they are again attracted, and thus vibrate in this state of flux.

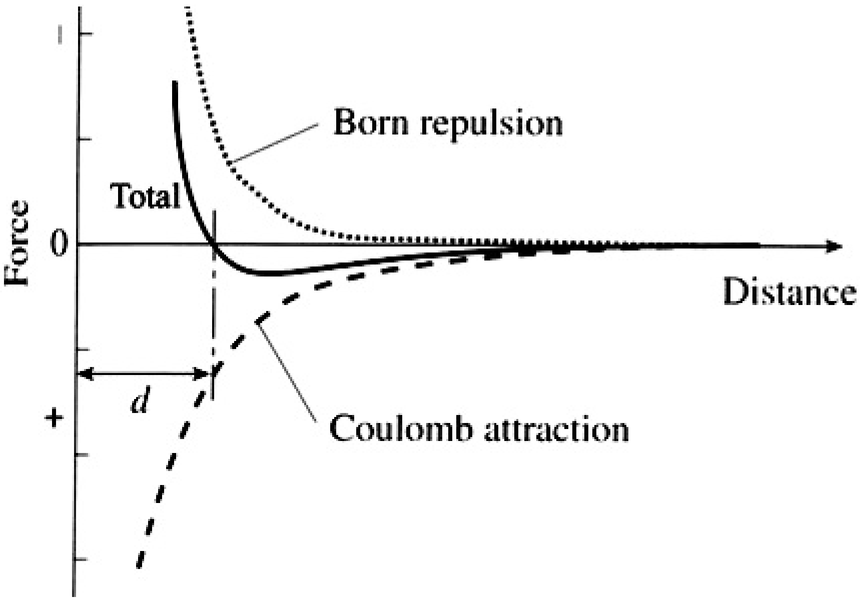


Figure 4.1: The repulsion of two ionically bonded atoms at different lengths

When forming a bond, some energy is released. This same amount of energy must again be added to the bonded compound to break those bonds. Some notes on the above diagram follow:

Bond length the distance between two bonded atoms at their minimum potential energy. Or, the average distance between two bonded atoms. This is the local min of the above graph.

Bond energy the amount of energy required to break a chemical bond to form neutral isolated atoms. This is the local max at the right-hand side of the graph.

4.5.1 Bonding of Electrons

The *octet rule* states that chemical compounds will tend to form such that each atom, by either gaining, losing, or sharing electrons has an octet of electrons in its highest occupied energy level.

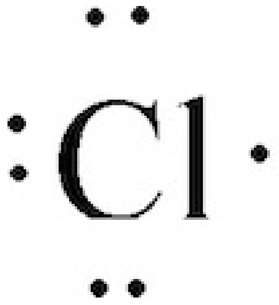
Exceptions

- Hydrogen does not form an octet, it only has two electrons around it.
- Boron & aluminum do not form octets, they have a total of 6 electrons surrounding them.
- Some p-block elements have more than 8 electrons.

4.6 Representation

4.6.1 Electron-Dot Structure

The *electron-dot structure* shows the amount of valence electrons around an atom. You only draw the dots of the valence electrons, and they are drawn around the chemical symbol of that element.



4.6.2 Lewis Dot Structure

The Lewis Dot Structure hinges on this idea of electron-dot structures, but extends it to molecules. Lewis Structures are formulae in which atomic symbols represent nuclei and inner-shell electrons. Dot-pairs or dashes between two atomic symbols represents the electron pair in bonds, and dots adjacent to only one atomic symbol represent unshared electrons.

Lone pair a pair of electrons that is not involved in bonding and belongs to only one atom

Structural formulae indicates the kind, number, arrangement, and bonds but not the lone pairs of atoms in a molecule.

To generate a Lewis Dot Structure, take the following steps:

1. Determine the type and number of atoms in the molecule
2. Write the electron-dot structure for each atom in the molecule
3. Determine the total number of valence electrons in the atom to be combined
4. Arrange the atoms to form a skeletal structure for the molecule. *Carbon is always the central atom in these structure if present, but otherwise pick the least-electronegative atom, and not hydrogen*
5. Connect the atoms by electron-pair bonds

6. Add unshared pairs of electrons so that each hydrogen atom shares a pair of electrons and each nonmetal has a full octet

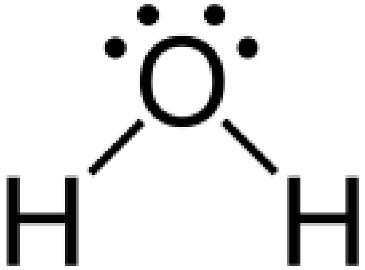


Figure 4.2: A Lewis dot-structure representation of a *Chlorine* ion

4.7 Bonding and Bonds

Single bond a covalent bond produced by the sharing of one pair of electrons between two atoms

Double bond a covalent bond produced by the sharing of two pairs of electrons between two atoms

Triple bond a covalent bond produced by the sharing of three pairs of electrons between two atoms

4.7.1 Resonance Structures

Sometimes, the Lewis Dot Structure does not accurately represent a molecule or ion. For example, in the case of O_3 *(ozone)*, several different structures can occur. This is called resonance, and cannot be captured by traditional Lewis structures.

Since the double bond can occur in multiple places (but only one place at a time) the Lewis structure is not suited to correctly display this. In the example below, the double-bond in the polyatomic ion *nitrate* is shown in multiple places throughout the structure.

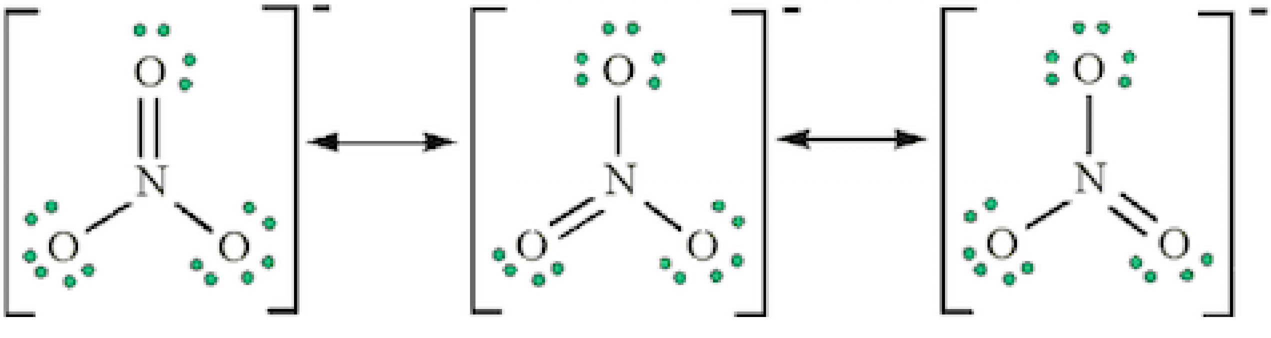


Figure 4.3: The double bond in this ion is placed in multiple locations throughout the diagram, to show that it can exist in more than one spot.

4.7.2 Metallic Bonds

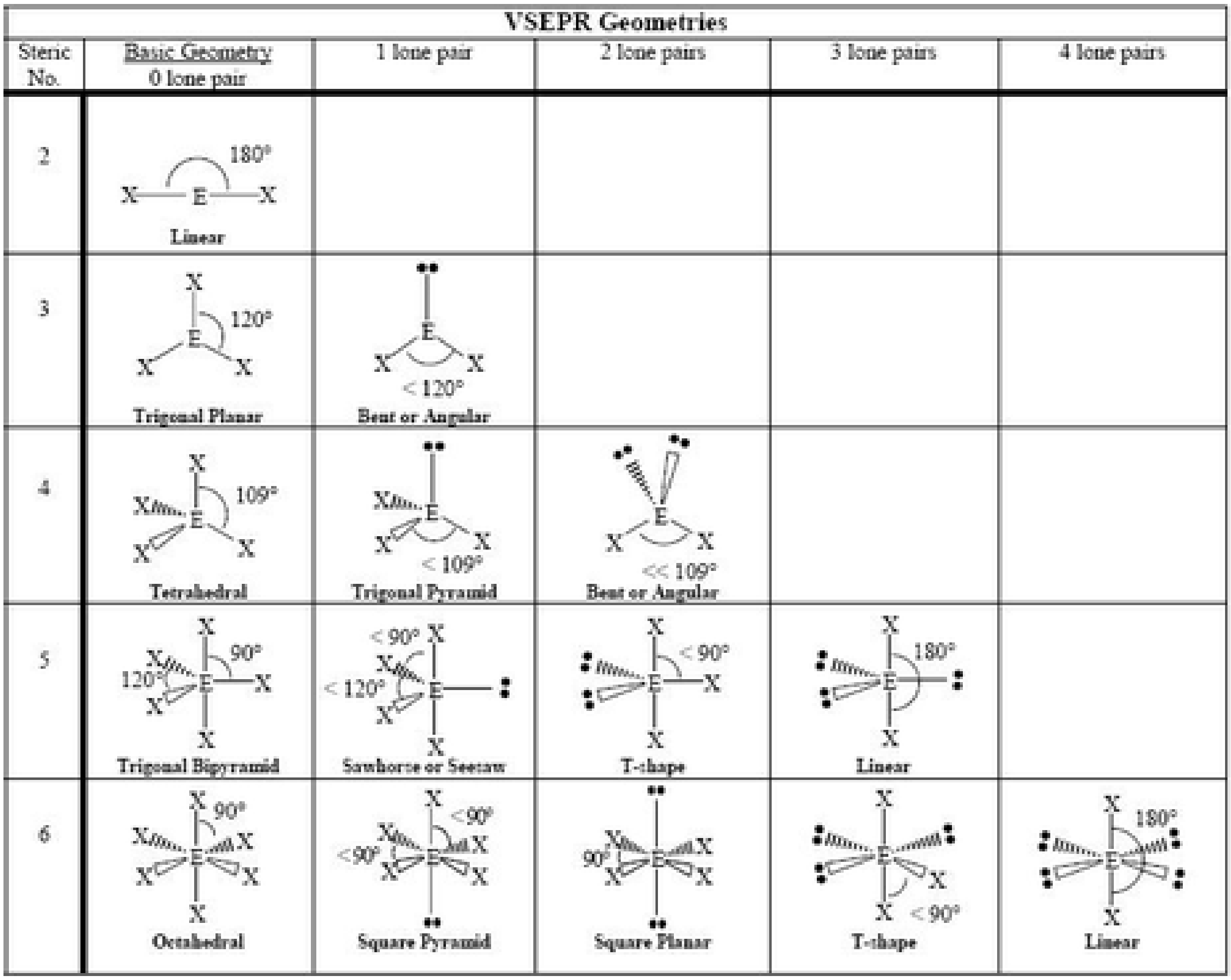
A *metallic bond* is the result of the attraction between metal atoms and the surroudning sea of electrons. It is acountable for most of the unique behavior of metals.

4.8 Molecular Geometry

While Lewis structures are all well and good, their major drawback is that they are only 2-dimensional. Molecular geometry is important because we can determine the molecular polarity or uneven distibution of molecular charge through structures in 3-dimensions.

4.8.1 VSEPR Theory

VSPER Theory states that the repulsion between sets of valence-level electrons surrounding an atom causes these sets to be oriented as far apart as possible.



VSEPR Geometries

Steric No.	Basic Geometry 0 lone pair	1 lone pair	2 lone pairs	3 lone pairs	4 lone pairs
2	180° Linear				
3	120° Trigonal Planar	< 120° Bent or Angular			
4	109° Tetrahedral	< 109° Trigonal Pyramid	<< 109° Bent or Angular		
5	90° 120° Trigonal Bipyramid	< 90° < 120° Sawhorse or Seesaw	< 90° T-shape	180° Linear	
6	90° Octahedral	<90° <90° Square Pyramid	90° Square Planar	< 90° T-shape	180° Linear

Chapter 5

Mass Relations and Stoichiometry

5.1 Atomic Mass

To determine the weight of an atom, scientists developed a small unit at which such quantities may be measured. This unit is know as the AMU, or *atomic mass unit.*

12 AMUs are defined as the weight of one atom of the Carbon-12 isotope. That is to say that $1 amu$ is exactly the weight of 1/12th of one atom of Carbon-12.

5.1.1 Masses of Elements

However, the numbers on the periodic table are not so round? Why might this be? Well, since we are accounting for isotopes, we must apply a weighted average to all naturally occuring isotopes. For example:

	Abundance	*Mass (amu)*
oxygen-16	99.762%	15.994
oxygen-17	0.038%	16.999
oxygen-18	0.200%	17.999

The calculation for the relative *amu* of Oxygen is as follows: $(0.99762 * 15.994) + (0.00038 * 16.999) + (0.00200 * 17.999) = 15.9994 amu$

5.2 The Mole

Avagadro's Number (or N_A) is thought of a quantity, usually of atoms. The number of that quantity is $6.022e23$. Therefore, since this is a quantity, the unit is $anything/mole$.

5.2.1 Molar Mass

Of Elements

To calculate the molar mass of an element: simply look to the periodic table! In the upper-left of the element box, you will find the relative molar mass of the element. That is to say, the *amu* of an element is the weight in grams of one mole of that element. For example, $1.008 mol$ of H weighs 1 g.

Of Compounds

Simply put, the molar mass of a compound is the sum of all its parts, multiplied individiually by each of the relative quantities.

5.3 Percent Composition

Percent composition of an element is the percent by mass of itself within a compound. For example, the molar mass of Cu_2S is 159.2. The mass amount of Cu

5.3.1 Empirical Formulae

An *emperical formula* consists of the sumbols for the elements combined in a compound, with subscripts showing the smallest whole-number mole ratio of the different atoms in the compound.

For example, the emperical formula P_2O_5 might lead to a chemical formula of P_4O_{10}, or similarily, the emp. formula C_3H_6O may lead to the formula $C_6H_{18}O_2$.

Calculating

To calculate an emperical formula, you must first start with the percentages of an atom or the mass of the atom present. Once these are determined, you may then convert to moles, and simplify the ratio by dividing by the smallest amount and rounding. If you cannot round cleanly, multiple the whole ratio by a factor.

The steps for calculating this emperical formula follow below:

1. Assume 100gs was sampled, unless given other measurements.
2. Convert the grams to moles, unless given moles.
3. Simplify the ratio of moles by dividing by the smallest term.
4. Multiply by factors to arrive at cleaner ratios, if not already clean.

5.4 Stoichiometry

Stoichiometry involves the mass relationships between reactants and products in a chemical reaction. IT allows you to determine the amount of additional reactants or products given the chemical equation and just one amount.

Stoichiometry relies on the ratios of moles between different compounds in a chemical reaction. Consider the following reaction: $Al_2(SO_4)_3 + 3\,Ca(OH)_2 \longrightarrow 2\,Al(OH)_3 + 3\,CaSO_4$. It has the following mass relations:

$$\frac{1 mol Al_2(SO_4)_3}{3 mol Ca(OH)_2}, \frac{1 mol Al_2(SO_4)_3}{2 mol Al(OH)_3}, \frac{1 mol Al_2(SO_4)_3}{3 mol CaSO_4}$$

To work with these mass relations, you can convert from mass → molar mass → moles → mole-to-mole ratio → moles → molar mass → mass.

5.4.1 Limiting Reagent

Sometimes when doing stoichometry, one reagent appears in a limiting amount such that you cannot fully complete the reaction. These substances are known as *limiting reagents*. The substances that is not completely used up in a reaction is referred to as an *excess reagent (reactant)*.

To determine the limiting reagent, just do the reaction with both against one of the products and figure out which makes *less*. The excess reagent will be what is left over. To determine the quantity of the excess reagent, use the quantity of the limiting reagent that you originally plugged in and wire that through.

Percent Yield Calculations

Sometimes, you are given an amount of which is supposed to occur (via stoichiometry) and an amount that you actually produced. To calculate this, you use the following:

$$100 * \frac{actual}{theoretical}$$

Chapter 6

Chemical Equations and Reactions

6.1 Introduction

Chemical reaction the process by which one or more substances are changed into one or more different substances.

Chemical equation a representation of a chemical reaction with symbols and formulae, the identies and relative amounts of reactants and products in a chemical reaction. An example follows: $(NH_4)_2Cr_2O_7 \longrightarrow N_2 + Cr_2O_3 + 4\,H_2O$

6.2 Indications

1. Evolution of heat and/or light energy
2. Production of a gas
3. Formation of a percipitate
4. Change in color

6.3 Characteristics of a Chemical Equation

1. The equation must represent known facts
2. The equation must contain the correct formulae for the reactants and products
3. The law of conservation of mass must be satisfied

- The coefficients represent relative, not absolute amounts of the substances in the equation.
- The relative masses can be determined as such (see chapter on molar mass)
- The reverse chemical equation holds the same properties and descriptors as specified above.

6.4 Balancing

Since the law of conservation of mass must be obeyed throughout a chemical reaction, the two sides of the equation must be *balanced*. That is to say, that the amount of atoms of an element on one side *must* equal the amount of atoms of that same element on the other side of the equation.

The steps to balancing a chemical equation follow below:

1. Write a formula equation by substituting the correct formulae for the reactants and products (only if a ”skeleton equation is not present)
2. Balance the formula equation according to the Law of Conservation of Mass:
   (a) Balance different types of atoms one at a time
   (b) First balance atoms of elements that are combined and only appear once on each side
   (c) Balance polyatomic ions that appear in both sides in a single unit *as a group*
   (d) Balance the H and O elements last.
3. Ensure that the atoms are balanced

6.5 Types of Reactions

Synthesis Reaction occurs when two or more substances combine to form a new compound.

Decomposition Reaction occurs whena single compound undergoes a reaction that produces two or more simpler substances

Single-replacement reaction occurs when one element replaces a similar element in a compound

Double-replacement reaction occurs when ions of two compounds exchange places, usually in the presence of an aqeuous solution, to form two new compounds

Combustion Reaction occurs when a substances combines with O_2 to release a large amount of energy in the form of light and heat

6.5.1 Synthesis Reaction

A synthesis reaction occurs when two or more substances make one substance. The reaction-type is as follows: $A + B \longrightarrow AB$.

6.5.2 Decomposition Reaction

This is the opposite of a synthesis reaction, that is to say that one compound breaks down into its simpler substances, usually with the input of heat or electricity. The reaction-type is as follows: $AB \longrightarrow A + B$.

6.5.3 Single-replacement Reaction

A single-replacement reaction takes one element and a compound, and swaps the similar elements: $A + BC \longrightarrow C + AB$.

If a metal is present, the positive element in the metal is swapped, if negaitve, otherwise (?).

6.5.4 Double-replacement Reaction