**Megathread**

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## Part 1: Atomic Structure

Isotopes – atoms of the same element (same atomic number) with a different number of neutrons (resulting in different nucleon numbers)

Isobars – atoms of different elements (different atomic numbers) with the same nucleon number

### Bohr’s Atomic Model

Rutherford’s model stated that an atom consisted of a nucleus, which is positively charged and is surrounded by electrons, which are negatively charged and move around in predictable paths called orbits.

Bohr stated that the electrons move around in fixed orbital shells, and that each shell had a fixed energy level. Electrons that have more energy are on orbitals that are further away from the nucleus, while electrons that have less energy are on orbitals that are closer to the nucleus.

### Postulates of Bohr’s Atomic Model

* Electrons move around the nucleus in fixed circular paths called orbits
* When electrons revolve in a particular orbit, they do not transfer energy. Such orbits are called stationary orbits.
* If an electron jumps from one orbit to another, it releases or gains energy as required in the form of radiation.
* Each orbit can hold up to electrons, where is the orbit number.

### Existence of Spectrum

When an electron jumps from a higher energy orbital to a lower energy orbital, it emits energy in the form of radiation. The radiation emitting from a particular material can be diffracted to reveal the specific frequencies of emitted radiation. This is called the emission spectrum.

Similarly, when energy is supplied to a material in the form of radiation, specific frequencies will be absorbed as some electrons jump from lower energy orbitals to higher energy ones. This radiation can then be diffracted to reveal that only those specific frequencies are missing. This is called the absorption spectrum.

Absorption and emission spectra can help determine what a material is. This is due to the fact that the atoms of different materials have orbitals with different energy levels. Thus, the specific frequencies of radiation their electrons will absorb are also different and unique.

### Quantum Numbers

Quantum numbers are a set of values that describe the state of an electron including its distance from the nucleus, the orientation and type of orbital where it is likely to be found, and its spin.

There are 4 quantum numbers, the principle quantum number (), the orbital angular momentum quantum number (), the magnetic quantum number () and the electron spin quantum number ().

### AUFBAU Principle

The AUFBAU principle states that in the ground state of an atom or ion, electrons fill atomic orbitals of the lowest available energy levels before occupying higher levels. Thus, the order goes: 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 6s, 5f, 6d and so on.

## Part 2: Derivations

### Derivation of Radius

Let an electron of charge revolve around a nucleus of charge , where is the atomic number.

Coulomb’s Force =

Centrifugal Force =

For hydrogen, .

From Bohr’s postulates we know,

The values of , and are known. Thus,

where is the principle quantum number.

For hydrogen,

This is known as the Bohr radius. It was assumed by Bohr to be the radius of a hydrogen atom in ground state.

### Derivation of Energy of an Electron in Orbit

Total Energy = Kinetic Energy + Potential Energy

### Derivation of Change in Energy During Orbital Change

(Rhydberg Constant) =

### Effect of Temperature on Solubility

From the 1st and 2nd Laws of Thermodynamics,

For gas and liquid solutions, the symbols and are used instead.For solid and liquid solutions, the symbols and are used instead.

For endothermic reactions, increases, so increases, so increases.

For exothermic reactions, increases, decreases (since is negative) and decreases.

### Rault’s Law

According to Rault’s Laws,

Here, represents the mole fraction, represents the partial pressure of the solvent, represents the partial pressure of the solution, represents the moles of solute and represents the moles of solvent.

This can now be used to find the molecular weight.

### Osmotic Pressure

### Boiling Point Elevation

In dilute solutions,

From Rault’s Law,

Here, is a constant so it can be removed and is in kilograms and must be converted to grams.

Here, is a constant known as the boiling point constant.

### De Broglie

### Schrodinger’s Wave Equation

Considering 3 dimensions,

## Part 3: Solutions

Solution – homogenous mixture of two or more substances meaning solute is evenly distributed throughout solvent; substances may be solids, liquids or gasses.

### Types of Solutions

Gas + Gas – Air

Gas + Liquid – Carbonated Drinks

Liquid + Liquid – Ethanol + Water (miscible)

Solid + Liquid – NaCl + Water Keep stirring and solid dissolves. At saturation point, for a certain temperature, solid stops dissolving. If temperature is increased, more solid can dissolve.

Solid + Solid – Gold + Silver Alloy

For solutions of liquids and gasses or liquids and solids, the liquid is the solvent. In every other case, the substance present in a smaller amount is considered the solvent.

### Equilibrium Vapor Pressure

The pressure created by the vapor phase of a substance on its liquid phase, when the rates of evaporation and condensation have reached an equilibrium.

### Boiling Point

The point at which the vapor pressure of a liquid is equal to the pressure in the environment around the liquid.

### Units of Concentration

Molarity – moles of solute per liter solution

Molality – moles of solute per kg solvent

Normality – grams of solute per liter solution

Formality – molecular weight of solute in grams per liter solution

Percentage – solvent/solution X 100%

Molarity, Normality and Formality affected by temperature.

Moles = weight/molecular weight

### Colligative Properties of Dilute Solutions

Colligative properties of solutions are properties that depend upon the concentration of solute molecules or ions, but not upon the identity of the solute. They are used to find the mass of an unknown solute.

* Lowering Vapor Pressure: Vapor pressure of a solvent decreases when a non-volatile solute is dissolved in it (related to Rault’s Law)
* Elevating Boiling Point: Boiling point of a solvent increases when a non-volatile solute is dissolved in it. The solute molecules will take up surface area, leaving less available space from which solvent molecules can vaporize, thus decreasing vapor pressure. Lower vapor pressure, means an increased boiling point.

### Raoult’s Law of Ideal Solution

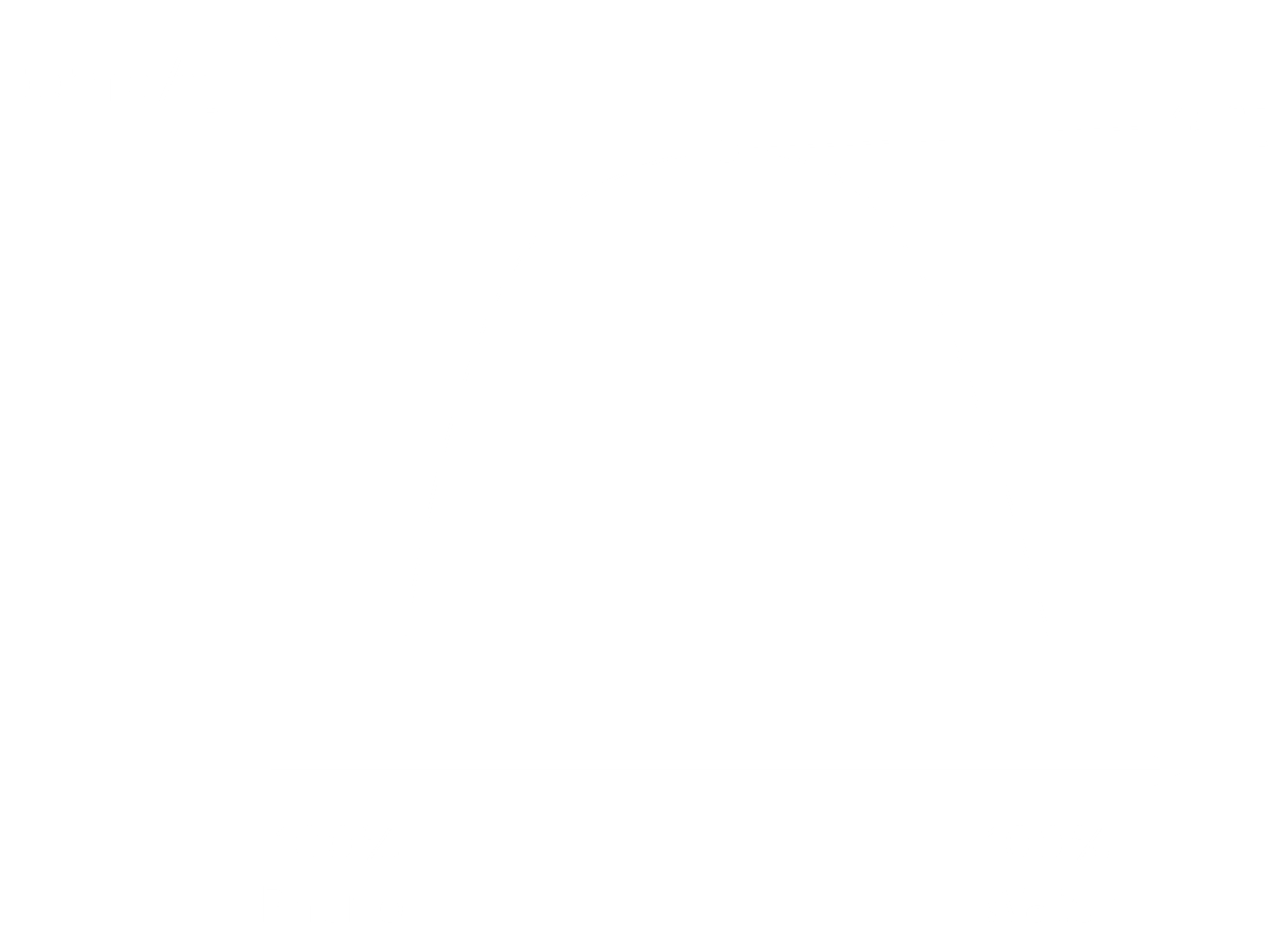
States that the vapor pressure of a solution is dependent on the mole fraction of a solute added to solution.

In an ideal solution, all substances present abide by Rault’s Law. Thus, the total vapor pressure is given by the partial vapor pressures of each substance.

### Critical Solution Temperature

Temperature above or below which two immiscible liquids become miscible for all proportions. For some liquids this occurs above a certain temperature (upper CST) and for some it occurs below a certain temperature (lower CST).

For Water + Phenol solutions,



For Water + Triethylamine solutions,



### Henry's Law

The amount of dissolved gas in a liquid is proportional to its partial pressure above the liquid at constant temperature.

### Applications of Henry’s Law

Used to increase solubility of carbon dioxide in carbonated drinks.

### Osmosis

A process by which molecules of a solvent tend to pass through a semi-permeable membrane from a less concentrated solution into a more concentrated one.

### Osmotic Pressure

Pressure that, when applied to the solution, just stops osmosis.

### Reverse Osmosis

A hydrostatic pressure greater than the osmotic pressure is applied in order to force a solvent to pass through a porous membrane in the direction opposite to that for natural osmosis.

### Usage of Reverse Osmosis

Reverse Osmosis is used for water purification. The solute containing ions, unwanted molecules and larger particles is retained on the pressurized side, while the pure solvent passes through to the other side.

### Van’t Hoff’s Theory of Osmotic Pressure

Van’t Hoff found that the solute particles in dilute solutions possess kinetic energy and move in random directions in the solutions, thus having behavior similar to that of gas molecules. On collision against a semipermeable membrane, the solute molecules exert osmotic pressure equal to the pressure which the solute molecules would exert if it were a gas molecule at the same temperature and occupying the same volume as that of the solution. Thus, the gas laws are applicable to dilute solutions.

At constant temperature, for molecules with osmotic pressure at volume.

At constant volume, .

This can be combined to

It was also proven that , thus for molecules,

## Part 4: Chemical Bonding and Molecular Orbital Theory

Chemical Bond – attraction between atoms allowing formation of chemical substances containing more than one atom; bond caused by electromagnetic force of attraction between opposite charges or dipole-dipole attraction

Ionic Compounds – Positive and negative ions; high melting/boiling point; conducts electricity in solution; dissolves easily

Covalent Compounds – Electrons shared; low melting/boiling points; poor electrical conductor;

Metallic Bonding – positive ions in sea of electrons; electrons move freely; strong bonds; malleable; conducts heat and electicity

Hydrogen Bonding – atoms with , or atoms; dipole-dipole attraction

Isomers – Same general formula different structures causing different properties

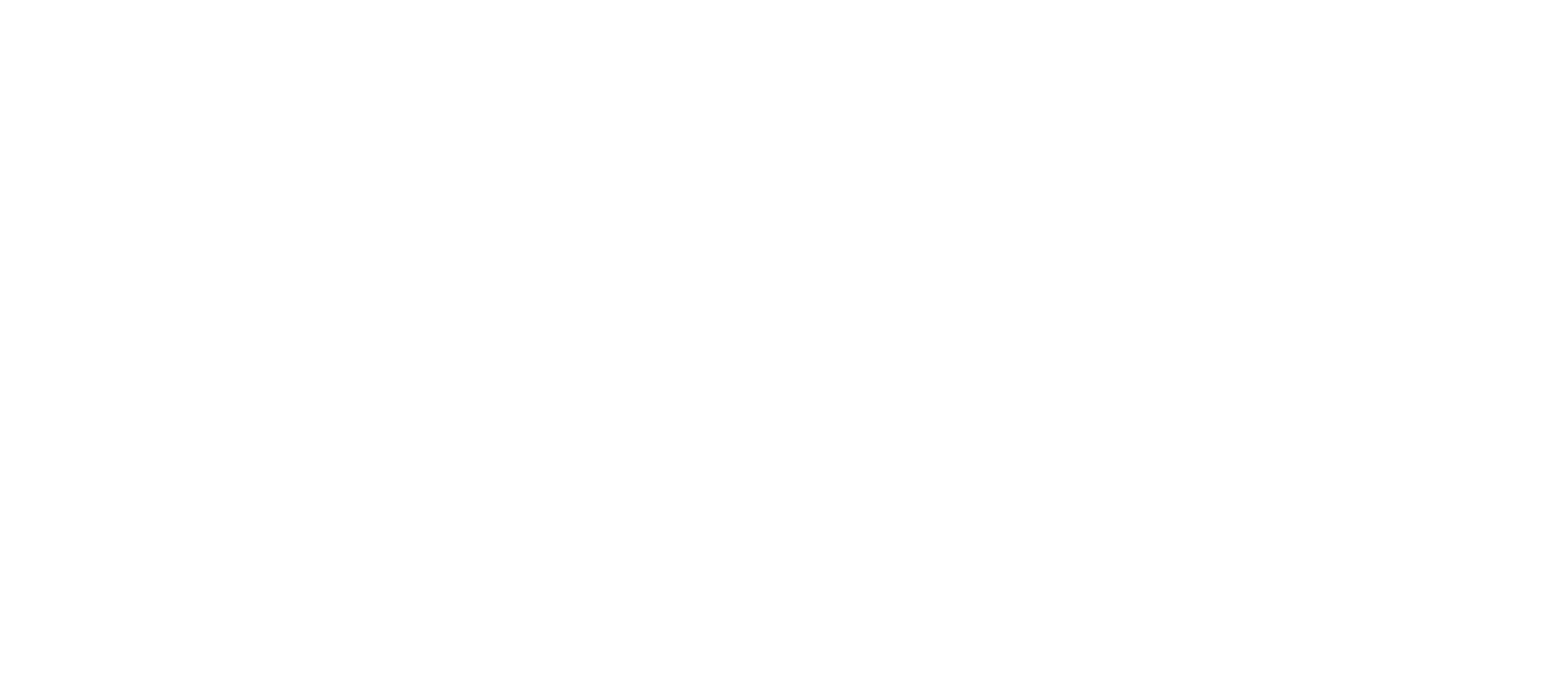
Electronegativity – ability to attract electrons; smaller and more protons means more electronegative

Partial Charges – electrons shift towards more electronegative atom in a bond; causes dipole moment

Lewis Dot Structures – dots are valence shell electrons; lines are bonds;

Molecular Orbital Theory – more electrons in area between nuclei where bond is formed – bonding orbitals; no electrons between two nuclei – anti-bonding; non-bonding make no difference – not valence shell;

Bond Order - ; always ceil; disregard non-bonding orbitals



Electronic Configuration for molecule, bonding order

Hund’s Rule – Every orbital on a single sublevel (, and ) must be filled with one electron before a second can be added; orbitals with single electrons will have same spin to maximize total spin

Paramagnets – subshells with unpaired electrons; weakly attracted by magnetic fields

Diamagnets – no unpaired electrons; repelled by magnetic fields

Periodic Table – Groups – Columns; Rows – Periods; Ionic Radii – increases down a groups, decreases from left to right for cations and increases left to right for anions; Ionization Energy – Decreases down a group, increases left to right; Metallic Character – Increases down a group, decreases left to right; Electronegativity – Decreases down a group, increases left to right

## Part 5: Chemical Kinetics

Order – Number of elements actively taking part in the reaction

Molecularity – number of elements taking part in the reaction

### Law of Mass Action

For ,

rate

### First Order Reactions

rate

– kinetic equation for a first order reaction in differential form

- kinetic equation for a first order reaction in linear form

- kinetic equation for a first order reaction in exponential form

- kinetic equation for a first order reaction in integral form

### Tests for First Order Reactions

* Same value of rate constant for different initial concentrations under same conditions
* Graph is straight line
* Constant half-life

### Second Order Reactions

rate

- kinetic equation for a second order reaction in differential form

- kinetic equation for a second order reaction in linear form

- kinetic equation for a second order reaction in integral form

### Tests for Second Order Reactions

* Isolation Method – Concentration of all components except one greatly increased so their effect is nullified

### Activation Energy Graphs

### Relationship Between Temperature and Rate of Reaction

– Arrhenius Parameter

increases, increases and vice versa

Exothermic – decreases, decreases; endothermic – decreases, increases

Le Chatelier’s Principle – If dynamic equilibrium is disturbed by changing conditions, the position of the equilibrium shifts to counteract that change. Affected by pressure, volume, concentration, catalyst and temperature.

### Equilibrium

Using Law of Mass Action,

Since or ,

## Part 6: Thermochemistry

Heat of Reaction – change in enthalpy of a chemical reaction at constant pressure

Heat of Solution – change in enthalpy when a substance is dissolved in a solution

Heat of Formation – change in enthalpy during formation of 1 mole of substance from its constituent elements, with everything in standard state

Heat of Combustion – change in enthalpy when one mole of a substance is completely oxidized

Heat of Neutralization – change in enthalpy when a gram equivalent of an acid and a gram equivalent of a base under neutralization

Hess’s Law – If a chemical equation can take place in different ways, following different sequences, then the overall enthalpy change for the reaction is the same regardless of which sequence was followed; can be used to find heat of formation/neutralization etc., heat of reaction for reactions that are slow or do not take place directly; intrinsic energy of an element is

### Formulae

where is the heat of reaction at constant pressure and is the heat of reaction of constant volume

where

where is the heat capacity at constant volume

where is the heat capacity at constant pressure

where is the standard heat of formation

Heat of neutralization = + heat of disassociation of weak acid/base

If is negative, exothermic (releases energy). If is positive, endothermic (absorbs energy).

Calorimeter – measures heat given out or absorbed during reaction

Water Calorimeter – insulated water bath; measure temperature rise;

Bomb Calorimeter – measures heat of combustion; similar; combustion takes place inside steel bomb with oxygen

### Gibb’s Free Energy

For moles of a substance ,

Where

Thus,

## Part 7: Electrochemistry

Electrolyte – forms ions in solution which conduct electric current;

Electrolytic Conductance – the transport of electric charges under electric potential difference by charges particles (ions) of atomic size or larger

Transport number – amount of total electricity transferred by cations and anions

Galvanic Cell – contain electrods with two different elements and their solutions

Salt Bridge – used to transfer ions; made of ; completes circuit