

Basic Chemistry

(CHEM1007.4)

Marmara University School of Engineering

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Note: Wx stands for xth week's lesson.

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Basic Chemistry (CHEM1007.4)

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Translations of Terms

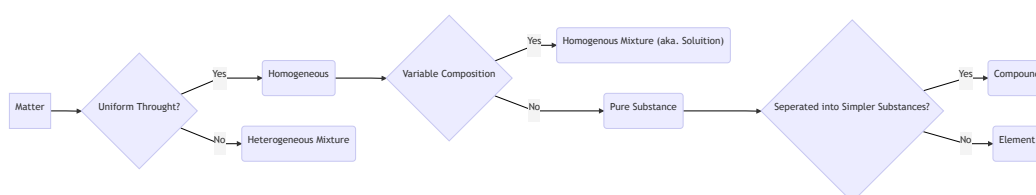
English	Turkish
Compound	Bileşik
Intensive Properties	Maddeye özgü özellikler
Extensive Properties	Maddeye özgü olmayan özellikler
Derived Units	Türetilmiş Büyüklükler
Substance	Madde
Flask	Şişe
Accuracy	Doğruluk
Precision	Hassaslık
Solubility	Çözünürlük
Spontaneous	Kendiliğinden olan
Limiting Reactant	Sınırlayıcı Bileşen
Combustion Reactions	Yanma Tepkimeleri
Yield	Verim
Electron promotion	Elektron atlaması (üst seviyeye)
Electron demotion	Elektron düşmesi (alt seviyeye)
Solute	Çözünen
Solvent	Çözücü
Restrain	Ölçülemek, Tutmak, Sınırlamak
Saturated	Doymuş
Semipermeable	Yarı geçirgen
Crenation	Plazmoliz
Hemolysis	Deplazmoliz
Buffer Solutions	Tampon Çözeltiler
Solubility Product	Çözünürlük Çarpımı
Precipitation Reactions	Çökeltme Tepkimeleri
Metathesis Reactions	Yer Değiştirme Tepkimeleri

W1 | Introduction: Matter and Measurement

- Chemistry is a science which is interested in matter and their properties. Matter is defined as anything that has a mass and takes up space.

Type of Matter	Description
Atoms of an element	Only one kind of atom in any element.
Molecules of an element	Only one kind of atom in any element.
Molecules of a compound	Must have at least two kinds of atoms.
Mixture of elements and compounds	At least two types of elements or compounds.

- Classification of matters and how to find the type of any matter can be represented with this diagram:



- There are types of properties: Intensive and Extensive. **Intensive properties** are independent from the amount of the substance but **extensive** ones are dependent. Density, boiling point, colour are intensive; mass, volume, energy is extensive.

- There are two types of changes: Physical and Chemical. Physical change doesn't affect the composition of a substance but chemical changes does, they make new substances. Therefore chemical reaction means a reaction which creates new substances.

Separation Types	Mixtures or Description
Filtration	Solid substances from liquids or solutions
Distillation	Homogeneous mixture into its components with difference in their boiling points
Chromatography	Separation substances on the basis of differences in solubility in a solvent.

- SI units are these:

Physical Quantity	Abbreviation and Name of Unit
Mass	Kilogram --- kg
Length	Meter --- m
Time	Second --- sec or s
Temperature	Kelvin --- K
Amount of substance	Mole --- mol
Electric current	Ampere --- A or amp
Luminous intensity	Candela --- cd

- 1 liter is a cube 1 decimeter long on each side. So 1 liter = 1 decimeter cube.
- Celsius and Kelvin relation: $K = C^{\circ} + 273.15$
- **Significant figures** rules: (0.004004500 -> 40045 significant ones)
 - All nonzero digits are significant.
 - Zeroes between two significant figures are themselves significant.
 - Zeroes at the beginning of a number are never significant.
 - Zeroes at the end of a number significant if a decimal point is written in the number.
- When addition is performed, answers are rounded to the least significant decimal place. In multiplication, answer's digit number must be same as the number of digits in least element.

- **Accuracy** refers to the proximity of measurement to the true value of a quantity. **Precision** refers to the proximity of several measurements to each other.

W1 | Atoms, Molecules and Ions

- **Dalton's Postulates**

1. Each element is composed of extremely small particles called atoms.
2. All atoms of a given element are identical but the atoms of one element are different from the atoms of all other elements.
3. Atoms of one element cannot be changed into atoms of a different element by chemical reactions; atoms are neither created nor destroyed in chemical reactions.
4. Compounds are formed when atoms of more than one element combine; a given compound always has the same relative number and kind of atoms.

- **Law of Conservation of Mass:** The total mass before the reaction is the same as the total mass after the reaction.

- **Law of Constant Composition:** In a compound, the relative numbers and kinds of atoms are constant.

- **Law of Multiple Proportions:** If two elements A and B combine to form more than one compound, the masses of B can combine with a given mass of A are in the ratio of small whole numbers.

- Electrons are discovered in 1897 by J. J. Thomson. Streams of negatively charged particles were found to emanate from cathode tubes, causing fluorescence then named as electrons.

- Cathode is the negative side (leaving side) and anode is the positive side (entering side).

- Thomson measured the charge/mass ratio the electron to be $1.76 \times 10^8 \frac{C}{g}$

- **Millikan's Oil-Drop Experiment:** Once the charge/mass ratio of the electron was known, determination of either the charge or the mass of an electron would yield the other. This experiments has found the charge of electron as $1.602 \times 10^{-19} C$

- In 1896, Henri Becquerel discovered that a compound of uranium spontaneous emits high-energy radiation. **Radioactivity** is the spontaneous emission of radiation of an atom.

- There are three types of radiation which are discovered by Rutherford:

- α particles: He^{2+}
- β particles: A high-energy, high-speed electron or positron emitted by the radioactive decay of an atomic nucleus during the

process of beta decay. There are two forms of beta decay, β^- decay and β^+ decay, which produce electrons and positrons respectively.

- γ rays

Atom Models	
Thomson's Model	<ul style="list-style-type: none"> * Seeds in a watermelon * Positive charge spread throughout sphere * Negative electrons embedded in sphere * Rutherford and Marsden proved this model wrong
Rutherford's Model	<ul style="list-style-type: none"> * A particle gun experiment (golden foil exp.) * Discovery and first seen of dense nucleus * A atom contains lots of empty space

- After that proton has discovered by Rutherford and neutron has discovered by Chadwick.
- Protons and electrons are the only particles that have a charge. Protons and neutron nearly has the same mass but electrons are nothing besides them.
- **Isotope** means same elements with different masses because of neutrons.
- Atomic mass unit is extremely useful when we're dealing with really low masses. 1 atomic mass unit is equals to mass of the unbound ^{12}C atom.
- **Average mass** is calculated from the isotopes of an element weighted by their relative abundances.

$$\text{average weight} = \sum \left(\text{isotope} * \frac{\text{isotope abundance}}{\text{abundance of the element}} \right) \quad (1)$$

- There are two types of formulas: **empirical formula** and **molecular formula**. Empirical formulas give the lowest whole-number ratio of atoms of each element in a compound. Molecular formulas give the exact number of atoms of each element in a compound. Also, structural formulas show the order in which atoms are bonded but perspective drawings also show the three-dimensional array of atoms in a compound.
- When atoms lose or gain electrons they become **ions**:
 - If it gains electrons it'll be anode.
 - If it lose electrons it'll be cathode.

- Ionic compounds can be made by ionic bonds.
- **Oxidation number:** the electron number which an atom gives or takes when become compound. Rules for determining oxidation numbers:
 - Oxidation number is zero if it's only element -- free atom. Ex.
 $\text{Cl} \rightarrow 0, \text{Al} \rightarrow 0$
 - Oxidation number is zero if it's neutral species summation. Ex.
 $\text{CH}_3\text{OH} \rightarrow 0, \text{MgCl}_2 \rightarrow 0$
 - Oxidation number is equals to charge of ion if it's an ion. Ex.
 $\text{Fe}^{3+} \rightarrow +3, \text{MnO}_4^- \rightarrow -1$
 - Group 1 metals have oxidation number +1 and group 2 metals have +2 at their compounds. Ex.
 $\text{K}_2\text{CO}_3 \rightarrow \text{K}^{+1}, \text{MgBr}^2 \rightarrow \text{Mg}^{+2}$
 - Fluorine at its compounds has a oxidation number of -1. Ex.
 $\text{SF}_6 \rightarrow \text{F}^{-1}$
 - Hydrogen at its compounds oxidation number is +1. Ex.
 $\text{NH}_3 \rightarrow \text{H}^{+1}$
 - Oxygen at its compounds oxidation number is -2. Ex. $\text{H}_2\text{O}, \text{CO}_2$

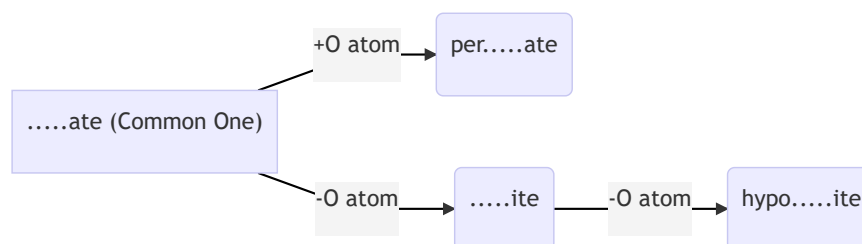
- **Inorganic Nomenclature**

1. Write the name of the cation.
2. If the anion is an element, change its ending to *-ide*; if the anion is a polyatomic ion, simply write the name of the polyatomic ion.
3. If the cation can have more than one possible charge, write the charges as a Roman numeral in parentheses.
4. Binary non-metal compound examples:

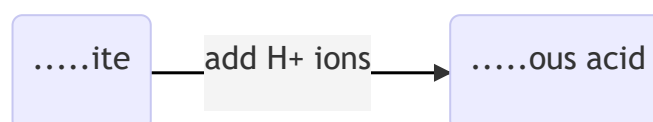
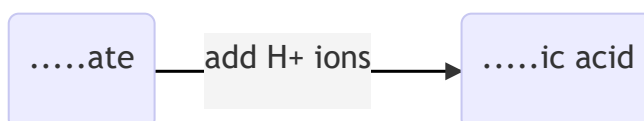
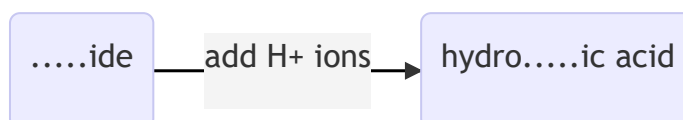
Formula	Name
HCl	Hydrogen chloride
SO ₂	Sulphur dioxide
SO ₃	Sulphur trioxide
B ₂ Br ₄	Dibor tetrabromide
NO ₂	Nitrogen dioxide
N ₂ O ₃	Dinitrogen trioxide

5. Binary acid examples. HCl : Hydrochloric acid, HBr : Hydrobromic acid

6. When there are two oxyanions involving the same element; the one with fewer oxygens ends in *-ite*, the one with more oxygens ends in *-ate*.



7. If the anion in the acid ends in *-ide*, change the ending to *-ic acid* and add the prefix *hydro-*.



8. If the prefix ends with *a* or *o* and the name of the element begins with a vowel, the two successive vowels are often elided into one.
Ex. N_2O_5 : Dinitrogen pentoxide

Prefix	Meaning
Mono	1
Di	2
Tri	3
Tetra	4
Penta	5
Hexa	6

Prefix	Meaning
Hepta	7
Octa	8
Nona	9
Deca	10

- Organic chemistry has its own nomenclature.
 - Alkanes: Only single bonds between hydrogen and carbon.
Formula: $C_x H_{4x}$
 - Alcohol: A hydrogen in an alkane is replaced with OH. An alcohol ends in *-ol*. Formula: $C_x H_{4x-1} OH$

W2 | Stoichiometry: Calculations with Chemical Formulas and Equations

- **Law of Conversion of Mass** is described by Antonie Lavoisier. Nothing is created and amount of matter exists both before and after the experiments is equals.
- **Limiting reactant** determines end of the reaction; when the limiting reactant finishes the reaction is finished. It has the smallest stoichiometric amount.

Reaction type	Description	Example
Combination Reactions	Two or more substances react to form <i>one</i> product.	$\text{N}_2 (\text{g}) + 3 \text{H}_2 (\text{g}) \longrightarrow 2 \text{NH}_3 (\text{g})$
Decomposition Reactions	One substance breaks down into two or more substances.	$2 \text{NaN}_3 (\text{s}) \longrightarrow 2 \text{Na} (\text{s}) + 3 \text{N}_2 (\text{g})$
Combustion Reactions	Rapid reactions that produce a flame which is often involved hydrocarbons reacting with oxygens in the air.	$\text{CH}_4 (\text{g}) + 2 \text{O}_2 (\text{g}) \longrightarrow \text{CO}_2 (\text{g}) + 2 \text{H}_2\text{O} (\text{g})$

- Formula weight is the mass of the elements with their amounts in formula. For example, CaCl_2 's formula weight is 110.99 amu which can be seen in this equation:

$$1 * Ca + 2 * Cl = 1 * 40.08 \text{ amu} + 2 * 35.453 \text{ amu} = 110.99 \text{ amu} \quad (2)$$

- Molecular weight is the mass of the elements with their amounts in molecule. It's different from formula weight because C_2H_6 's formula weight is CH_3 's mass but its molecular weight isn't.
- **Percent Composition** is the mass ratio of a particular element in a compound.

$$\% \text{ Element} = \frac{(\text{number of atoms}) * (\text{atomic weight})}{\text{formula weight of the compound}} * 100 \quad (3)$$

- Mole is a unit that shows the number of things in a packet of $6.02 * 10^{23}$ things.
- **Combustion Analysis**
 - C is determined from the mass of CO_2 produced.
 - H is determined from the mass of H_2O produced.
 - O is determined by difference after the C and H have been determined.
- Theoretical yield is the amount of the product possible as calculated through the stoichiometry problem. Actual yield is the amount one actually produces and measures. Percent yield is the ratio of actual yield and theoretical yield in percent.

W3 | Electronic Structure of Atoms

- **Wavelength** is the distance between corresponding points on adjacent waves.
- The number of waves passing a given point per unit of time is the **frequency**.
- The relationship between frequency and wavelength is this:

$$c = f * \lambda \quad (4)$$

- Quanta means a packet of energy. It's explained by Max Planck. Einstein used this to explain and understand photoelectric effect.
- Planck Equation is below which h is Planck constant, $6.626 * 10^{-34} Js$

$$E = h * f = h * \frac{c}{\lambda} \quad (5)$$

- **Monochromatic** means single wavelength radiation; **polychromatic** is combined wavelength radiations.
- Niels Bohr's postulates from Planck's assumption:
 1. Electrons in an atom can only occupy certain orbits (corresponding to certain energies).
 2. Electrons in permitted orbits have specific, "allowed" energies; these energies will not be radiated from the atom.
 3. Energy is only absorbed or emitted in such a way as to move an electron from one "allowed" energy state to another; the energy is defined by $E = h * v$
- The energy absorbed or emitted from the electron promotion/demotion can be calculated by this equation where $R_h = Rydberg\ constant = 1.097 * 10^7 m^{-1}$:

$$\Delta E = -hcR_h \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right) \quad (6)$$

- Heisenberg showed that the more precisely the momentum of a particle is known, the less precisely is its position known with **the uncertainty principle**:

$$(\Delta x) * (\Delta mv) \geq \frac{h}{4\pi} \quad (7)$$

- Quantum mechanics is developed with a mathematical equation for explaining both wave and particle nature of matter in an equation by Erwin Schrödinger. The wave equation is designed with ψ , square root of wave equation gives a probability density map where an electron be

able in any point. Solving the wave equation gives us wave functions which named as **orbitals** and their energies.

- **Quantum Numbers**

- *Principle Quantum Number* (n : $n \geq 1$): Describes the energy level on which the orbital resides.
- *Angular Momentum Quantum Number* (l : $n - 1 \geq l \geq 0$): The shape of the orbital.

Value of l	0	1	2	3
Type of orbital	s	p	d	f

- *Magnetic Quantum Number* (m_l : $l \geq m_l \geq -l$): Describes the the 3D orientation of the orbital. Therefore, there's 1 s orbital, 3 p orbitals, 5 d orbitals and 7 f orbitals.

n	l	Subshell	m_l	Orbital Count
1	0	1s	0	1
2	0	2s	0	
	1	2p	-1, 0, 1	4
3	0	3s	0	
	1	3p	-1, 0, 1	
	2	3d	-2, -1, 0, 1, 2	9
4	0	4s	0	
	1	4p	-1, 0, 1	
	2	4d	-2, -1, 0, 1, 2	
	3	4f	-3, -2, -1, 0, 1, 2	16

- s Orbital: Observing a graph of probabilities of finding an electron versus distance from the nucleus, we see that s orbitals possess $n - 1$ nodes, or regions where there is 0 probability of finding an electron.
- p Orbital: They have two lobes with a node between them.
- d Orbital: Four of the five d orbitals have 4 lobes; the other resembles a p orbital with a dough nut around the center.
- *Spin Quantum Number* (m_s : $-\frac{1}{2}$ or $+\frac{1}{2}$): The “spin” of an electron describes its magnetic field, which affects its energy.

- **Pauli Exclusion Principle:** There can't be two electrons in the same atom can have exactly the same energy (or quantum numbers).
- **Hund's Rule:** For degenerate orbitals, the lowest energy is attained when the number of electrons with the same spin is maximized. In understandable way, electrons in a subshell must be placed on the same spin then the other spin if there are electrons left.

W4 | Electronic Structure of Atoms

- When forming compounds, atoms tend to add or subtract electrons until they are surrounded by eight valence electrons -- it is known as the octet rule.
- There are three types of exceptions to the octet rule. These molecules or ions in the conditions below do not follow the rule.
 - Ions or molecules with an odd number of electrons,
 - Ions or molecules with less than an octet,
 - Ions or molecules with more than eight valence electrons
- Lattice energy is the energy required to completely separate a mole of a solid ionic compound into its gaseous ions. It increases with the charge on the ions and decreasing size of ions.
- In polar covalent bonds, the electrons are not always shared equally. For example, fluorine pulls harder on the electrons -- which is relevant with electronegativity. The greater the difference in electronegativity, the more polar is the bond.
- Electronegativity increases as you go
 - from left to right across a row,
 - from the bottom to the top of a column.
- **Dipole moment**, μ , produced by two equal but opposite charges separated by a distance, r , is calculated:

$$\mu [\text{debyes } D] = Qr \quad (8)$$

- When the electronegativity difference is greater than 2.0, the bond is considered ionic.
- Lewis structures are representations of molecules showing all electrons, bonding and nonbonding.
 - Find the sum of all valence electrons in the polyatomic ion or the molecule.
 - If it is an anion, add one electron for each negative charge.
 - If it is a cation, subtract one electron for each positive charge.
 - The central atom is the least electronegative element that isn't hydrogen. Connect it to the outer atoms with single bonds. Track of the remaining electrons.
 - Fill the octets of the outer atoms. Keep track of the electrons.
 - Fill the octet of the central atom. Keep track of the electrons.

- If you run out of electrons before the central atom has an octet form, multiple bonds until it does.
- Then assign the formal charge. Formal charge is the difference between the valence electrons and electrons assigned to atom.
- Resonance structures are the molecules or ions with changeable bonds.
- Bond enthalpy is the energy required to break the bond and also known as strength of a covalent bond.

$$\Delta H_{rxn} = \sum H_{broken} - H_{formed} \quad (9)$$

- As the number of bonds between two atoms increases, the bond length decreases.

W5 | Periodic Properties of the Elements

- Two years after Rutherford proposed the nuclear model of the atom, English physicist Henry Moseley developed the concept of atomic numbers.
- Mendeleev and Meyer independently came to the same conclusion about how elements should be grouped.
- The effective nuclear charge is the subtracting of atomic number and screening constant.

$$Z_{eff} = Z - S \quad (10)$$

- Bonding atomic radius is defined as 1/2 of the distance between covalently bonded nuclei. It decreases from left to right across a row and increases from top to bottom of a column.
- Ionic size depends on three factors: the nuclear charge, the number of electrons, the orbitals in which electrons reside.
 - Cations are smaller than their natural atoms.
 - Anions are larger than their natural atoms.
 - Ions size increases from top to bottom and from right to left. The snowman hypothesis. :-)
- Ionization energy is the amount of energy required to remove an electron from the ground state of a gaseous atom or ion.
 - x'th ionization energy means energy required to remove x'th electron.
 - As one goes down a column, less energy is required to remove the first electron.
 - As one goes across a row, it gets harder to remove an electron. However, the "3 aşağı 5 yukarı" rule must be considered.
- Electron affinity is the energy change accompanying the addition of an electron to a gaseous atom.
 - In general, electron affinity becomes more exothermic as you go from left to right across a row.
- Metallic character increases with right to left and top to bottom.

Metals	Nonmetals
Shiny Luster	Don't have a luster

Metals	Nonmetals
Malleable and Ductile	Brittle, Hard or Soft
Good conductors	Poor conductors
Metal oxides are ionic solids are basic	Most non-metal oxides are molecular substances form acids
Form cations in aqueous	Form anions or oxyanions in aqueous

- Alkali metals are soft and metallic solids.
 - They are found only in compounds in nature.
 - They have low densities, melting points and low ionization energies.
 - Their reactions with water are famously exothermic.
 - They react with oxygen to form peroxides.
 - K, Rb and Cs also form super-oxides.
 - They produce bright colors when placed in a flame.
- Alkaline Earth Metals,
 - They have higher densities and melting points than alkalis.
 - Their ionization energies are low but not that much as alkalis.
 - Reactivity increases as you go up to down in the group.
- In Group 6A, oxygen, sulfur and selenium are non metals. Tellurium is a metalloid. The radioactive polonium is a metal.
- In Group 7A, we call them as Halogens (Greek, tr. Salt Formers), they have large and negative electron affinities. They react with metals to form metal halides.
- In Group 8A, the noble gases, have astronomical ionization energies. Their electron affinities are positive, so they are nonreactive. They are found as monatomic gases.

W6 | Gases

NEED TO UPDATE OR REWRITE FOR MORE INFORMATION!

- Atmospheric pressure is the weight of air per unit of area. $P = F/A$
- Unit Pascal is equals to 1 N/m^2 . Also, 1 bar (unit) is equals to 10^5 Pa .
- But, besides that, we measure the atmospheric pressure with *atm* or *torr*.
- Standart Pressure = $1.00 \text{ atm} = 760 \text{ mmHg} = 101.325 \text{ kPa}$
- Boyle's Law

$$V \propto \frac{1}{P} \quad (11)$$

- Charles's Law

$$V \propto T_{abs} \quad (12)$$

- Avogadro's Law

$$V \propto n \quad (13)$$

- So, the ideal gas equations that we can derive from the previous equations is,

$$PV = nRT \quad (14)$$

- Density of a gas from the ideal gas equation and $d = \frac{m}{v}$

$$d = \frac{m}{V} = \frac{PM}{RT}, \quad M = \text{molecular mass} \quad (15)$$

- Effusion Speed:

$$u_{rms} = \sqrt{\frac{3RT}{M}} \quad (16)$$

- Real gases -> high temp. and low press.
- Corrected Ideal Gas Equation is the Van der Waals Equation given below.

$$\left(P + \frac{n^2 a}{V^2}\right)(V - nb) = nRT \quad (17)$$

W7 | Liquids and Intermolecular Forces

NEED TO UPDATE OR REWRITE FOR MORE INFORMATION!

- London Dispersion Forces
 - At one instant time, the electrons move into one position -- which they doesn't do it regularly because of repulsion between same charged particles, and that makes the particle dipole.
 - If there's another induced dipole, then the polarity of the particles make them connected with a force which we called as London Dispersion.
 - Larger surface area increases the dispersion force.
 - Molecular weight increases the dispersion force.
- Dipole-Dipole Interactions
 - Molecules that have permanent dipoles are attracted to each other.
 - If the molecules nearly have same properties, the dipole-dipole interactions will be much bigger than the London dispersion force.
- Hydrogen Bonding is a particular type of Dipole-Dipole interactions between H and N, O or F. It is usually a weak force.
- Ion-Dipole Interactions
 - This force is related with ionic substances dissolving in polar solvents.
- Phase Changes
 - Solid -> Liquid: Melting/Fusion
 - Liquid -> Gas: Vaporization
 - Solid -> Gas: Sublimation
 - Liquid -> Solid: Freezing
 - Gas -> Liquid: Condensation
 - Gas -> Solid: Deposition
- Liquid molecules evaporate and vapor molecules condense at the same time.
- Clausius - Clapeyron equation is shown below.

$$\ln P = -\frac{\Delta H_{vap}}{RT} + C, \quad c \text{ is a constant} \quad (18)$$

W7 | Solids and Modern Materials

NEED TO UPDATE OR REWRITE FOR MORE INFORMATION!

- **Primitive cubic lattice's** lattice points only at its corners.
- **Body-centred cubic lattice's** lattice points at corners and the centre of the unit cell.
- **Face-centred cubic lattice's** lattice points at the corners and the centre of the each surface of its unit cell.
- Covalent-network solids tend to be hard and have high melting points.
- If you add 3A, it will become p-type. If you add 5A, it'll become n-type.
- Vulcanization, chains are cross-linked by short chains of sulphur atoms, making the rubber stronger and less susceptible to degradation.

W10 | Properties of Solutions

- Solutions are homogenous (uniform) mixtures. **Solute** dispersed throughout the solvent.
- **Entropy** rises when the molecules mix and become randomly distributed.
- In order to form a solution, the intermolecular forces between the solute and solvent particles must be strong enough to compete with those between solute particles (the particles in the solute) and those between solvent particles (the particles inside the solvent).
- The solvent pulls the solute particles apart and surrounds them to become a solution.
- The enthalpy change of the overall process can be written as below. In the equation, variable as follows:
 - $\Delta H_{\text{solution}}$: Enthalpy change for solution
 - $\Delta H_{\text{solvent}}$: Enthalpy change for separation of solvent particles.
 - ΔH_{solute} : Enthalpy change for separation of solute particles.
 - ΔH_{mix} : Enthalpy change to create new interactions between solute and solvent.

$$\Delta H_{\text{solution}} = \Delta H_{\text{solvent}} + \Delta H_{\text{solute}} + \Delta H_{\text{mix}} \quad (19)$$

- For exothermic processes, the $\Delta H_{\text{solution}}$ will be less than zero.
 - For endothermic processes, the $\Delta H_{\text{solution}}$ will be greater than zero.
- Dissolution is a physical change — you can get back the original solute by evaporating the solvent. If you can not get it back, it means that it has reacted.
- Types of solutions are saturated ones, unsaturated ones and supersaturated ones. In the saturated solutions, dissolved solute is in dynamic equilibrium with solid state particles. On the other hand, in supersaturated solutions, the solvent holds more solute than is normally possible at that temperature.
- **Like dissolves like!** It means that polar substances tend to dissolve in polar solvents and nonpolar tend to dissolve in nonpolar solvents.
- The more similar the intermolecular attractions, the more likely one substance is to be soluble in another.
- Larger molecules have stronger dispersion forces, so that in general, solubility of gases in water *increases* with increasing mass.

- Solubility of a gas in a liquid is directly proportional to its pressure.
- Solubility of the gas, with **Henry's Law**:

$$S_g = kP_g \quad (20)$$

- The relation between temperature and solid's solubility in liquid solvents is proportional. But the relationship between temperature and gas's solubility in liquid solvents is inversely proportional.
- Carbonated soft drinks are more "bubbly" if stored in the refrigerator.
- To change molarity to molality or vice versa, we need the density of the solution.
- Colligative properties depend only on the number of solute particles present.

- Vapor Pressure

- Higher concentrations of non-volatile solutes make it harder for solvent to escape to the vapor phase.
- **Raoult's Law**: X_A is the mole fraction of compound A, and P_A° is the normal vapor pressure of A at that temperature.

$$P_A = X_A P_A^\circ \quad (21)$$

- Boiling Point Elevation and Freezing Point Depression

- Non-volatile solute and solvent interactions cause solutions to have **higher boiling point** and **lower freezing points** than the pure solvent.
- K_b is the molal boiling point elevation constant. i is the **van't Hoff** factor.

$$\Delta T_b = K_b * m * i \quad (22)$$

$$T_{sol} = T_{pure} + \Delta T_b \quad (23)$$

- K_f is the molal freezing point depression constant. i is the **van't Hoff** factor.

$$\Delta T_f = -K_f * m * i \quad (24)$$

$$T_{sol} = T_{pure} + \Delta T_f \quad (25)$$

- As you see, ΔT does not depend on what type the material is, but how many particles are dissolved.
- When solute particles associate in solution, $i < 1$.
- When solute particles dissociate in solution, $i > 1$.
- When solute particles neither dissociate nor associate in solution, $i = 1$.

- Osmosis

- In biological systems, most semipermeable membranes allow water to pass through but solutes are not free to do so.
- The net movement of solvent from higher solvent concentration to lower solvent concentration. In the equation above, π known as osmotic pressure, not π from the mathematics. M is the molarity of the solution.

$$\pi = \left(\frac{n}{V} \right) RT = MRT \quad (26)$$

- *Hypertonic* means the solute concentration outside the cell is greater than the inside. *Hypotonic* is the opposite one.
 - When the water will flow out of the cell the **crenation** results.
 - Water will flow into the cell and **hemolysis** results.
- Suspensions of particles larger than individual ions or molecules, but too small to be settled out by gravity are called **colloids**.
 - Colloidal suspensions can scatter rays of lights and this phenomenon is known as the **Tyndall effect**.

W11 | Chemical Kinetics

- The more homogeneous the mixture of reactants, the faster the molecules can react.
- As concentration of reactants increases, likelihood that reactant molecules will collide much more.
- Since at higher temperatures, reactants will move faster; collide will be more often or with greater energy.
- Catalysts speeds up the reactions but they are not consumed during the reaction.
- We can find the rate of reactions by monitoring the change in concentration.
 - As the reaction proceeds, the average rate decreases. It's because there are fewer collisions left at the end.
- The relation between reactions rates and stoichiometry is wrote bellow.



$$\text{Rate} = -\frac{1}{a} \frac{\Delta[\text{A}]}{\Delta t} = -\frac{1}{b} \frac{\Delta[\text{B}]}{\Delta t} = \frac{1}{c} \frac{\Delta[\text{C}]}{\Delta t} = \frac{1}{d} \frac{\Delta[\text{D}]}{\Delta t} \quad (28)$$

- The exponents tell the order of the reaction with respect to each reactant. For the expression bellow, the individual orders for NH_4^+ and NO_2^- is first order. But the reaction of this rate is second-order overall.

$$\text{Rate} = k[\text{NH}_4^+][\text{NO}_2^-] \quad (29)$$

- Let us integrate the rate law to find $[\text{A}]_t$'s function. If you look deeper to the *equation 33*, you can see that it's form of linear function graphs (see *equation 34*). Therefore, we can say that for a first-order reaction the plot of $\ln[\text{A}]_t$ vs t will yield as a straight line with slope of $-k$.

$$\ln \left(\frac{[\text{A}]_t}{[\text{A}]_0} \right) = -kt \quad (30)$$

$$\ln[\text{A}]_t - \ln[\text{A}]_0 = -kt \quad (31)$$

$$\ln[\text{A}]_t = -kt + \ln[\text{A}]_0 \rightarrow y = ax + b \quad (32)$$

- For second-order processes, integrating the rate law in reactant A, we get *equation 34*, and it is also in the form of linear functions. So, the plot of $1/[\text{A}]_t$ vs t also yields a straight line with slope of k .

$$\frac{1}{[\text{A}]_t} = kt + \frac{1}{[\text{A}]_0} \quad (33)$$

- **Half-life** defined as the time required for one-half of a reactant to react. For a first-order process, it does not depend on $[A]_0$.

$$[A]_t = 0.5[A]_0 \quad (34)$$

$$\ln \frac{0.5[A]_0}{[A]_0} = -kt_{1/2} \quad (35)$$

$$\ln(0.5) = -0.693 = -kt_{1/2} \quad (36)$$

$$t_{1/2} = \frac{0.693}{k} \quad (37)$$

- But for second order processes, it depends on $[A]_0$.

$$t_{1/2} = \frac{1}{k[A]_0} \quad (38)$$

- k is temperature dependent. As temperature increases, so does the reaction rate.
- Molecules have to collide in correct and specific orientation with enough energy.
- **Activation energy** is the minimum amount of energy required for reaction. In the reaction coordinate diagrams, the peak point is called as *transition state* and the species present at the transition state is called *activated complex*. The energy gap between the reactants and the activated complex is the *activation-energy barrier*.
- *Maxwell-Boltzmann Distributions* tells us that as the temperature increases, the curve flattens and broadens. Thus, at higher temperatures, a larger population of molecules has higher energy. In the *equation 40*, R is the gas constant and T is the absolute temperature. f is the fraction of molecules.

$$f = e^{-E_a/RT} \quad (39)$$

- Arrhenius Equation is given *equation 41*. A is the frequency factor. If you take natural logarithm both of its sides, the equation becomes linear function for $\ln(k)$ vs $\frac{1}{T}$. It means that if k is determined experimentally at several temperatures, E_a can be calculated from the slope of the graph.

$$k = Ae^{-E_a/RT} \quad (40)$$

- The table below clearly indicates what does **molecularity** means and what is the rate laws for different reactions.

Molecularity	Elementary Reactions	Rate Law
Uni molecular	$A \longrightarrow \text{products}$	$Rate = k[A]$
Bi molecular	$A + A \longrightarrow \text{products}$	$Rate = k[A]^2$

Molecularity	Elementary Reactions	Rate Law
<i>Bi</i> molecular	$A + B \longrightarrow \text{products}$	$Rate = k[A][B]$
<i>Ter</i> molecular	$A + A + A \longrightarrow \text{products}$	$Rate = k[A]^3$
<i>Ter</i> molecular	$A + A + B \longrightarrow \text{products}$	$Rate = k[A]^2[B]$
<i>Ter</i> molecular	$A + B + C \longrightarrow \text{products}$	$Rate = k[A][B][C]$

- In multistep processes, one of the steps will be slower than all others. That step will be the reaction to write rate law. Note that, termolecular processes are rare, so that if there's a termolecular processes, look for two-step mechanism.
- Catalysts increase the rate of a reaction by decreasing the activation energy of the reaction. Enzymes are the catalysts in biological systems.

W12 | Chemical Equilibrium

- Equilibrium means that the forward reaction speed is equal to the backward reaction speed. If we use rate law and make some algebraic manipulations, the $\frac{K_{forward}}{K_{backward}}$ becomes $K_{equilibrium}$ and can be formulated bellow. For $a A + b B \rightarrow c C + d D$

$$K_{eq} = \frac{[C]^c [D]^d}{[A]^a [B]^b} \quad (41)$$

- Since pressure is related to concentration for gases, we can write the equilibrium constant's formula as like shown bellow.

$$K_p = \frac{(P_C^c)(P_D^d)}{(P_A^a)(P_B^b)} \quad (42)$$

- From the Ideal Gas equation, we can derive K_p from the K_c . Here, δn means the subtraction of moles of gaseous reactant from the moles of gaseous product.
($\delta n = (\text{moles of gaseous product}) - (\text{moles of gaseous reactant})$)

$$K_p = K_c (RT)^{\delta n} \quad (43)$$

- Equilibrium can be reached from either direction. It doesn't matter whether we start with reactants or products.
- If $K \gg 1$, the reaction is product-favoured; product predominates at equilibrium.
- If $K \ll 1$, the reaction is reactant-favoured; reactant predominates at equilibrium.
- Manipulating equilibrium constants with algebra has rules.
 - If the reaction turned vice-versa, reactants became products and products became reactants, the constant will be reciprocated. It means that getting the $\frac{1}{K_{eq}}$.
 - If the reaction multiplied with a number, the constant will be powered with the number. (For multiplying with 2, constant will be K_{eq}^2 .)
 - The equilibrium constant for a net reaction made up of two or more steps is the product of the equilibrium constants for the individual steps.
- The concentrations of solids and liquids (not ions) do not appear in the equilibrium expressions.
- *The Reaction Quotient (Q)* gives the same ratio the equilibrium expression gives, but for a system that is not at equilibrium.

- If $Q < K$, reaction forms products because there is too much reactant.
- If $Q = K$, reaction is at equilibrium.
- If $Q > K$, reaction forms reactants because there is too much product.
- **Le Châtelier's Principle** states that "If a system at equilibrium is disturbed by a change in temperature, pressure, or the concentration of one of the components, the system will shift its equilibrium position so as to counteract the effect of the disturbance."
 - If a matter added to the equilibrium, the reaction wants to decrease the matter so that the equilibrium stays.
 - For endothermic reactions, if you heat the equilibrium, it will want to decrease temperature (spend heat/energy). If you cool the equilibrium, it will want to increase temperature.
- Catalysts increase the rate of both direction of reactions. Equilibrium is achieved faster with using it.

W13 | Acid-Base Equilibria

- Arrhenius perspective on acid and bases is,
 - An acid is a substance that, when dissolved in water, increases the concentration of hydrogen ions. (H^+)
 - A base is a substance that, when dissolved in water, increases the concentration of hydroxide ions. (OH^-)
- Bronsted-Lowry perspective on acid and bases is,
 - An acid is a proton donor. It must have a removable acidic proton.
 - A base is a proton acceptor. It must have a pair of nonbonding electrons.
 - If it has both of these two situations, it is an **amphiprotic**.
- Lewis acids are defined as electron-pair acceptors. Atoms with an empty valence orbital can be Lewis acids. Lewis bases are defined as electron-pair donors. (Note: Lewis bases can interact with things other than protons.)
- **Conjugate** comes from the Latin and meaning "to join together". An example of conjugate bases and acids are given. In the example, HNO_2 is an acid because it removes H^+ and its conjugate base is NO_2^- . H_2O is a base because it adds H^+ and its conjugate acid is H_3O^+ .



- In acid-base reactions, equilibrium favours the production of the weaker acid and base. i.e.; the stronger the reacting acid and base, the more complete the reaction. Think of a Bronsted acid-base reaction as a competition between the 2 bases in the system for protons. The stronger base "wins" and forces the equilibrium in the direction of the weaker acid and base. Below, the H_2O and ClO_4^- ions, both bases, are competing for protons. H_2O , the stronger of the two, "wins" and the equilibrium is shifted to the right in the direction of the weaker base, ClO_4^- ions. HClO_4 , the stronger acid, has a greater tendency to give its protons to H_2O than H_3O^+ ions have to give their protons to ClO_4^- ions. The equilibrium is displaced so far to the right that the reaction is essentially complete.



- Autoionization of water is an equilibrium reaction. The equilibrium constant is referred to as the **ion product constant** for water.



- pH is defined as the negative base-10 logarithm of the concentration of hydronium ion. $pH = -\log[H_3O^+]$. There are other "p" scales such as pOH and pK_w . $pOH = -\log[OH^-]$ and $pK_w = -\log K_w$.

$$pH + pOH = pK_w = 14.00 \quad (47)$$

- Some strong acids are HCl, HBr, HI, HNO₃, H₂SO₄, HClO₃, HClO₄ and some strong bases are Ca²⁺, Sr²⁺, Ba²⁺.
- The equilibrium constant for acid-dissociations is called the **acid-dissociation constant**, K_a . The greater K_a , the stronger is the acid. The equilibrium constant for base-dissociations is called the **base-dissociation constant**, K_b .
- Percent ionization is the ratio of last molarit to initial molarit of acid.

$$Ionization = \frac{[H_3O^+]_{eq}}{[HA]_{initial}} * 100 \quad (48)$$

- Polyprotic acids are specific acids that are capable of losing more than a single proton per molecule in acid-base reactions.
- K_a and K_b are related in this way:

$$K_a * K_b = K_w \quad (49)$$

- Anions are bases, so that, they can react with water in a hydrolysis reaction to form OH⁻ and the conjugate acid.



- Cations with acidic protons (like NH_4^+) will lower the pH of a solution. Most metal cations that are hydrated in solution also lower the pH of the solution.
- Greater charge and smaller size make a cation more acidic.
- Effect of Cations and Anions are described here.
 - An anion that is the conjugate base of a strong acid will not affect the pH.
 - An anion that is the conjugate base of a weak acid will increase the pH.
 - A cation that is the conjugate acid of a weak base will decrease the pH.
 - Cations of the strong Arrhenius bases will not affect the pH.
 - Other metal ions will cause a decrease in pH.
 - When a solution contains both a weak acid and a weak base, the affect on pH depends on the K_a and K_b values.
- Acidity increases from left to right across a row and from top to bottom down a group for HX_n combinations.

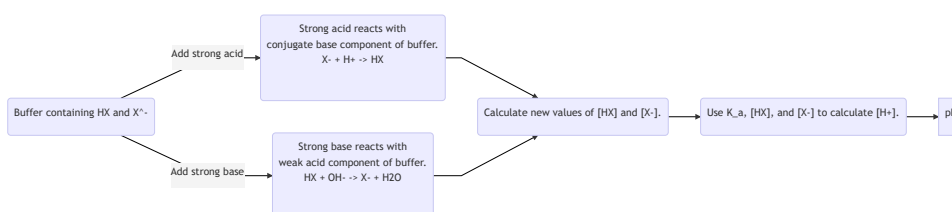
- In oxyacids, in which an -OH is bonded to another atom, Y, the more electronegative Y is, the more acidic the acid. For a series of oxyacids, acidity increases with the number of oxygens.
- Resonance in the conjugate bases of carboxylic acids stabilizes the base and makes the conjugate acid more acidic.

W13 | Acid-Base Equilibria

- The extent of ionization of a weak electrolyte is decreased by adding to the solution a strong electrolyte that has an ion in common with the weak electrolyte. For an example, see page 4 of the Chapter 17's presentation file.
- Buffers are solutions of a weak conjugate acid–base pair. We can calculate pH values of buffers with Henderson–Hasselbalch equation that can be seen below. They are particularly resistant to pH changes. For an example calculation, see page 16 of the Chapter 17's presentation file.

$$pH = pK_a + \log \frac{[base]}{[acid]} \quad (51)$$

- The pH range is the range of pH values over which a buffer system works effectively. It's best to choose an acid with a pK_a close to the desired pH.
- Solution process of addition a strong or base to a buffer is given schemetically as follows and described with texts.
 - Determine how the neutralization reaction affects the amounts of the weak acid and its conjugate base in solution.
 - Use the Henderson–Hasselbalch equation to determine the new pH of the solution.



- **Solubility product** (K_{sp}) is the rate of saturated solution. Don't confuse with solubility, these two are different things.



$$K_{sp} = [\text{Ba}^{2+}][\text{SO}_4^{2-}] \left[\frac{g}{L} \text{ or } \frac{g}{mL} \text{ or } M \right] \quad (53)$$

- The Common-Ion Effect: If one of the ions in a solution equilibrium is already dissolved in the solution, the equilibrium will shift to the left and the solubility of the salt will decrease.
 - pH: If a substance has a basic anion, it will be more soluble in an acidic solution. Substances with acidic cations are more soluble in basic solutions.
 - Complex Ions: Metal ions can act as Lewis acids and form complex ions with Lewis bases in the solvent. The formation of these complex ions increases the solubility of these salts.
- In a solution,
 - If $Q = K_{sp}$, the system is at equilibrium and the solution is saturated.
 - If $Q < K_{sp}$, more solid can dissolve.
 - If $Q > K_{sp}$, the salt will precipitate.

W13 | Reactions in Aqueous Solution

- Solutions are homogeneous mixtures. **Solvents** are the ones who solves the **solutes**.
- **Dissociation** is the process that the solvent pulls an individual ion from the crystal (ionic substance). Electrolyte is dissociates into its ions when dissolved in water. Strong electrolytes dissociates completely but weak electrolytes dissociates partially.
 - Strong acids are strong electrolytes.
 - Strong bases are strong electrolytes.
 - Soluble ionic salts are strong electrolytes.
- Soluble ionic compounds are generally electrolytes.
- Molecular compounds, on the other hand, are generally non electrolytes. Except the ones who is an acid or a base.
- The **molecular equation** lists the reactants and products in their molecular form. In the **ionic equation** all strong electrolytes (strong acids, strong bases, and soluble ionic salts) are dissociated into their ions. This more accurately reflects the species that are found in the reaction mixture.
- To form the **net ionic equation**, cross out anything that does not change from the left side of the equation to the right. Those things that didn't change (and were deleted from the net ionic equation) are called **spectator ions**.
 - Write a balanced molecular equation.
 - Dissociate all strong electrolytes.
 - Cross out all spectator ions.
 - Write the net ionic equations with remaining parts.
- Generally, when solutions of an acid and a base are combined, the products are a salt and water. It's called as **neutralization reactions**.
- Some metathesis reactions do not give the product expected.
 - When a carbonate or bicarbonate reacts with an acid, the products are a salt, carbon dioxide, and water.
 - Similarly, when a sulfite reacts with an acid, the products are a salt, sulfur dioxide, and water.
- An oxidation occurs when an atom or ion *loses* electrons.
- An reduction occurs when an atom or ion *gains* electrons.
- Oxidation number rules are,

- Elements in their elemental form has an oxidation number of 0.
- The oxidation number of a monatomic ion is the same as its charge.
- Nonmetals tend to have negative oxidation numbers, although some are positive in certain compounds or ions.
 - Oxygen has an oxidation number of -2 , except in the peroxide ion, in which it has an oxidation number of -1 .
 - Hydrogen is -1 when bonded to a metal, $+1$ when bonded to a nonmetal.
 - Fluorine always has an oxidation number of -1 .
 - The other halogens have an oxidation number of -1 when they are negative; they can have positive oxidation numbers, however, most notably in oxyanions.
- The sum of the oxidation numbers in a neutral compound is 0.
- The sum of the oxidation numbers in a polyatomic ion is the charge on the ion.
- In displacement reactions, ions oxidize an element. The ions, then, are reduced. The reverse reaction does not occur in this type of reactions because of their activities.
- Molarity is a way to measure the concentration of a solution.

$$\text{Molarity } [M] = \frac{\text{moles of solute}}{\text{volume of solution in liters}} \quad (54)$$

End of the story, for just now. Thank you for your contributions.