UNIT 1

Chem 30 Unit 1 MS1: Examine the role of valence electrons in the formation of chemical bonds.

Outline
- Democritus’ model
- Dalton's model
- J. J. Thomson's model
- Rutherford's model
- Light as waves + particles
- c = λf
- E = hf
- KE = hf – φ
- Spectra

- Bohr's model
- The quantum model
- Quantum numbers
- Orbital notation
- Electron configuration
- Aufbau principle
- Hund's rule
- Pauli exclusion principle
- Noble gas notation

- Periodic table and e- configuration
- Valence electrons
- Hadrons
- Isotopes
- Periodic trends: atomic radii, IE,
- Electron affinity + electronegativity
- Lewis structures
- VSEPR theory and geometry
- Molecular polarity

Democritus

Democritus, a Greek philosopher, is credited for first formalizing a particle theory of matter around 400 BC. He believed all matter to be made of small, indivisible particles he called atoms. (fr. Gk atomos, indivisible) His theory had 6 points:

- Atoms are particles in the void
- Atoms are always in ______
- Atoms are ____________
- Atoms come in different __________, shapes, and are arranged in different ways
- Atoms make life __________
- The human soul is made of __________

The idea of atoms was dismissed a generation later, by the more influential thinkers Plato and Aristotle. For almost 2200 years the concept of the atom was forgotten.

Dalton

John Dalton, a teacher from England, re-established an atomic theory in 1808, based on his experiments. He explained the law of conservation of mass using the laws of definite and multiple proportions. Dalton's atomic theory:

- All matter is composed of small particles called atoms.
- Atoms of the same element are __________, atoms from different elements differ in ______, ______ and other properties.
- Atoms cannot be __________, created, or destroyed.
- Atoms of different elements combine in small whole-number ratios to form __________.
- In chemical reactions, atoms are ______________, separated, or rearranged.

J. J. Thomson

By 1897, scientists such as William Crookes had been experimenting with ______________.

These were glass tubes with a gas at low pressure that glowed when subjected to high voltages. The experiments by many scientists, including Thomson and Crooke, found that the glow was made of particles streaming from the __________ terminal to the __________;

- The particles could __________ objects
- The particles were deflected by a __________ field
- The particles were deflected away from a __________ charged object

p. 65
p. 66
p. 70
Thomson was able to calculate a charge to mass ratio for these particles and found them to be the same for all materials. In 1903, he proposed his "plum pudding" model of the atom: atoms were made from positive material, laced with small negative particles. The negative particles were later named _____________. (electric + _on, fundamental particle)

Millikan
Robert Millikan used more accurate equipment in his oil drop experiment. He found the mass of an electron to be _______ times smaller than the smallest atom. (hydrogen) Here are the masses:

<table>
<thead>
<tr>
<th>Particle</th>
<th>Mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>proton (p⁺)</td>
<td>$1.673 \times 10^{-24}$ g</td>
</tr>
<tr>
<td>neutron (n⁰)</td>
<td>$1.675 \times 10^{-24}$ g</td>
</tr>
<tr>
<td>electron (e⁻)</td>
<td>$9.109 \times 10^{-28}$ g</td>
</tr>
</tbody>
</table>

these more accurate numbers were calculated later

Rutherford
Ernest Rutherford was a New Zealander who did some of his work in Canada. In 1911, He and some colleagues bombarded thin gold foil with _______ particles. The foil was about 400 atoms thick, and alpha particles are fast, positively charged particles with a mass of 4 times that of hydrogen. If the atom was uniformly dense, as was believed, the particles should have passed through with only small deflections.

This was the case, however, 1 in every 8 000 particles was __________ toward the source. Rutherford eventually concluded that the centre of an atom must be dense and __________, while most of the atom was empty space. He named the centre the __________, and didn't know where the electrons were, but the idea that they orbited the nucleus like planets orbit the sun was beginning to catch on. Protons in the nucleus should repel each other, but are held together by short-range attractive forces called __________ forces.

The _______ was later discovered by Irène and Frederic Joliot-Curie in 1932 and was named by James Chadwick.

Light as Waves and Particles
Before 1900, light was believed to behave as a wave. Waves have characteristics such as wavelength ( ), amplitude ( ), and frequency ( ). For all electromagnetic radiation, wavelength and frequency are inversely proportional in the equation

$$c = \text{the speed of light} \ 3.00 \times 10^8 \text{ m/s}$$
$$\lambda = \text{wavelength in m}$$
$$f = \text{frequency in Hz or cycles per second}$$

The wave model of light could not explain the different colours given off by glowing hot object like iron. ____________ figured out that atoms can gain or lose energy only in small, specific amounts called quanta. A ____________ is the smallest amount of energy that an atom can gain or lose. The wave model of light also could not explain the photoelectric effect.

Photoelectric Effect
When light is shined on a certain metal surface, electrons can be _________ and detected. However, if the ____________ of the light waves is too low, no electrons can be emitted. But if the frequency is high enough, even a single photon of light is enough to liberate an electron.
To explain the photoelectric effect, Einstein proposed in 1905 that light has properties of waves and particles. A _______ is a massless particle that carries a ______________ of energy. The energy of a photon can be calculated using Planck’s idea of the quantum or amount of energy.

\[
E = \text{the energy of a photon in J} \\
h = \text{Planck's constant} = 6.626 \times 10^{-34} \text{ Js} \\
f = \text{frequency in Hz}
\]

In other words, photons at a given frequency will possess a known amount of energy.

ex 1) Calculate the frequency of yellow light with a wavelength of 585 nm.

ex 2) Calculate the energy of a photon of x-ray radiation if its frequency is \(7.28 \times 10^{18}\) Hz.

\[
\text{Photoelectric effect calcs} \quad \text{where: KE = the max kinetic energy of the emitted electron in J} \\
h = \text{Planck's constant} = 6.626 \times 10^{-34} \text{ Js} \\
f = \text{frequency in Hz} \\
\phi = \text{binding energy of electrons in J}
\]

ex 1) Light with a wavelength of \(2.50 \times 10^{-7}\) m falls on a piece of chromium causing an electron to be emitted. What is the kinetic energy of the electron if the binding energy of electrons in chromium is \(7.21 \times 10^{-19}\) J?

ex 2) What is the longest wavelength possible that could emit a photon from chromium? (set KE to zero to find the limit where no electron will be emitted)

Experiments such as the double slit experiment showed wave properties of electrons.
Bohr

When hydrogen gas in a tube is energized using electricity, its electrons are _______ to a higher energy state. When they fall back to their ground state, they give off a _______ with energy equal to that gained. Its colour can be observed and calculated using \( E = hf \). When adding more electricity, we should see a continuous colour change, but in fact we only see __________.

In 1913, Niels Bohr proposed that electrons can occupy only certain __________. The electron orbital lowest in energy (ground state) is closest to the nucleus, and higher energy levels are successively further away. This model of the atom explained the hydrogen spectrum, but failed to account for many electron atoms. (see energy series fig. 4-8, 4-9, p.96)

The Quantum Mechanical Model de Broglie, Heisenberg, Schrödinger p. 98

In order to understand the quantum nature of the atom, scientists had to change their model of the electron. It had become accepted that light or energy could behave as both waves and particles, but in 1924, Louis de Broglie suggested that _______ could also behave as _______.

In order to detect where electrons are in the atom, if we use light or photons, they will knock an electron off course. Heisenberg's uncertainty principle states that the _______ certain you are of an electron's position, the _______ certain you are of its speed, and vice versa. So instead of talking about where electrons are, we use electron probability _________ or areas where you are likely to find an electron.

This dual nature of the electron led Erwin Schrödinger to develop the Schrödinger wave equation. This is the foundation for quantum theory.

\[
\frac{d^2 \psi}{dx^2} + \frac{8 \pi^2 m}{h^2} (E - V) \psi = 0
\]

where: \( \psi = \) Schrödinger wave function
\( x = \) position
\( h = \) Planck's constant
\( E = \) total energy
\( V = \) potential energy

Quantum Numbers

From this equation, we receive 4 quantum numbers to describe electrons and orbitals:
1. **Principle quantum number** (\( n \)): describes the __________ of the electron. It also determines how many orbitals you have. There are \( n^2 \) solutions to the Schrödinger equation, or \( n^2 \) orbitals. 

\[ n = 1, n = 2, n = 3... \]

2. **Angular momentum quantum number** (\( l \)): describes the __________ when \( l = s \) orbital 
   \[ l = 0 \text{ orbital} \]
   \[ l = 1 \text{ orbital} \]
   \[ l = 2 \text{ orbital} \]
   \[ l = 3 \text{ orbital} \]
   \[ l = 4 \text{ orbital} \]
   Note: these are also called subshells

3. **Magnetic quantum number** (\( m_l \)): describes the orientation or __________ of the orbital. (x, y, z) 
   It ranges from **-** to **-**. 

4. **Magnetic spin quantum number** (\( m_s \)): describes the __________ of the electron. If there are two electrons in the same orbital, they must spin in opposite directions: **-** or **-**. There can only be 2 electrons per orbital. Spin is also shown with arrows: **-** or **-**.

**Electron Configuration: Orbital Notation**  
According to the Aufbau principle, an electron will occupy the ______-energy that will accept it. (aufbau is German for "building up")

- energy level 1 (the lowest)

ex) hydrogen has 1 electron

\[ \text{1s} \]

-s orbital (first to fill)

ex 2) nitrogen has 7 electrons

\[ \begin{array}{cccc}
1s & 2s & 2p_x & 2p_y & 2p_z \\
\end{array} \]

-1s fills, then 2s, then the 2p (there are 3 of them)

-2 electrons in the same orbital have opposite spin ( vs. )

-according to Hund's rule, orbitals of equal energy levels are each occupied by _____ electron before 2 electrons are placed in the ________ orbital.

The filling order for orbitals is

\[ 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, 7p \]

with electrons:

ex 3) nickel has 28 electrons

\[ \begin{array}{cccccccccccccccc}
\uparrow \downarrow & \uparrow \downarrow & \uparrow \downarrow & \uparrow \downarrow & \uparrow \downarrow & \uparrow \downarrow & \uparrow \downarrow & \uparrow \downarrow & \uparrow \downarrow & \uparrow \downarrow & \uparrow \downarrow & \uparrow \downarrow & \uparrow \downarrow & \uparrow \downarrow & \uparrow \downarrow \\
1s & 2s & 2p & 3s & 3p & 3d & 4s & 4p & 5s & 5p & \text{6s} & \text{6p} & \text{7s} & \text{7p} \\
\end{array} \]

**Electron Configuration**

Electron configuration notation is a shorthand way of writing this:

or also

\[ 1s^22s^22p^63s^23p^63d^84s^2 \]

sometimes ordered according to energy level.
Periodic Table and Electron Configuration  

Understanding electron configuration brings new perspective to the periodic table. For example, there are 2 columns for the ________ elements because s orbitals can hold 2 electrons; there are 10 transition elements because d orbitals hold _______ electrons. Therefore:

Groups _____ are the s-block elements
Groups ______ are the d-block elements  
Groups ______ are the p-block elements
Lanthanides and Actinides are the _________ elements

When using the periodic table to determine electron configuration, note that the d block is on the energy level one lower than its period; the f block is 2 lower.

Noble Gas Notation  

Note that for nickel, the closest noble gas that has fewer electrons is Ar = 1s² 2s² 2p⁶ 3s² 3p⁶ so another way of writing nickel is

ex 2) iodine (53 e-) is

Pauli Exclusion Principle  

The Pauli exclusion principle states that _______ (in the same atom) have the __________ quantum numbers:

ex) neon (10 e-)

draw a line from the electron to the set of quantum numbers

Valence Electrons  

Valence electrons are the outermost s and p electrons. Although you elements will never have more than 8 valence electrons, (s² + p⁶) energy levels hold the following electrons:

<table>
<thead>
<tr>
<th>energy level</th>
<th>e⁻/shell</th>
<th>configuration</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>1s²</td>
<td></td>
</tr>
<tr>
<td>2</td>
<td>2s² 2p⁶</td>
<td></td>
</tr>
<tr>
<td>3</td>
<td>3s² 3p⁶ 3d¹⁰</td>
<td></td>
</tr>
<tr>
<td>4</td>
<td>4s² 4p⁶ 4d¹⁰ 4f¹⁴</td>
<td></td>
</tr>
<tr>
<td>5</td>
<td>5s² 5p⁶ 5d¹⁰ 5f¹⁴</td>
<td></td>
</tr>
<tr>
<td>6</td>
<td>6s² 6p⁶ 6d¹⁰</td>
<td></td>
</tr>
<tr>
<td>7</td>
<td>7s² 7p⁶</td>
<td></td>
</tr>
</tbody>
</table>

(valence)

Continued Discoveries  

Our understanding of the atom is always advancing. Recently, we have resolved our understanding of matter to include __________ – composite particles made from quarks held together by the strong force. There are two types of hadrons:

__________ – made from 3 quarks  
ex) protons and neutrons

__________ – made from 1 quark and 1 antiquark  
ex) pions and kaons
Isotopes and Average Atomic Masses

Isotopes are atoms of the ________ element that have different numbers of neutrons, and therefore have _________. For example, hydrogen

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Protons</th>
<th>Neutrons</th>
<th>Mass #</th>
<th>Abundance</th>
<th>Atomic Mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>Protium</td>
<td>1</td>
<td>0</td>
<td>1</td>
<td>99.985%</td>
<td>1.007825 amu</td>
</tr>
<tr>
<td>Deuterium</td>
<td>1</td>
<td>1</td>
<td>2</td>
<td>0.015%</td>
<td>2.014102 amu</td>
</tr>
<tr>
<td>Tritium</td>
<td>1</td>
<td>2</td>
<td>3</td>
<td>trace</td>
<td>3.016049 amu</td>
</tr>
</tbody>
</table>

To find the average atomic mass for hydrogen, take the atomic mass of each isotope and multiply it by its abundance (as a decimal) and add them together.

(In 1961, the average atomic mass of hydrogen was found to be 1.00797, but in 1981, this figure was revised to be _______ amu.)

The periodic table displays the average atomic masses for the elements in amu.
Ex) Boron has 2 natural isotopes. The mass of boron-10 is 10.01294 amu and the mass of boron-11 is 11.00931 amu. If the atomic mass of boron from the periodic table is 10.811 amu, calculate the relative abundance of the two isotopes to one decimal place.

The Periodic Table
Dmitri Mendeleev is regarded as the ________ of the periodic table. He noticed that when elements were arranged according to increasing atomic mass, similar chemical ________ recur at regular intervals. For example, Li, Na, and K have similar properties. His work led to the ________: when elements are arranged in order of increasing atomic number, elements with similar chemical and physical properties appear at regular intervals.

Review periodic table
Group 1: Alkali metals - very reactive, react with water to form base, low m.p.
  2: Alkaline earth metals - reactive, also form base in water, slightly higher m.p.
  3-12: Transition metals - a transition from reactive metals to nonmetals
  13: Boron family - metals and metalloid, not very reactive
  14: Carbon family - not very reactive 
  15: Nitrogen family - not very reactive
  16: Oxygen family - more reactive
  17: Fluorine family - very reactive, form halogenated compounds
  18: Noble gases - full valence shells, will not react under normal circumstances
  Lanthanides and Actinides: transition metals, elements 84+ are radioactive, elements 93+ are synthetic

Periodic Trends: Atomic Radii
The size of an atom (atomic radius) is determined by the ______ of the nucleus and the number of ______ levels. As we move down a group, we notice atoms ______ in size. This makes sense, because electrons are filling higher energy levels, which are further from the nucleus. Inner electrons also shield the outer electrons somewhat from the nucleus. However, as we move left to right across a period, atoms ______ in size. Although the number of electrons is increasing, they are on the same energy level, and are pulled closer to a more positive nucleus.

Periodic Trends: Ionization Energy
Ionization energy (IE) is the amount of energy required to ______ an electron, to form an ion. As we move down a group, we notice that IE ______. This is because the electrons are farther from the nucleus and shielded from the positive nucleus by the middle electrons. As we move across a period, IE generally ______ because there is a greater positive charge attracting each electron on the same energy level. (which is also closer to the nucleus) The first ionization energy (IE₁) refers to the energy required to remove the ______ electron from an
atom. To remove further electrons (IE₂, IE₃, etc.) is ______ difficult, as each electron is attracted to a more positive nucleus.

IE₁  IE₂  IE₃ ....

Periodic Trends: Electron Affinity
Electron affinity is the amount of energy ______ when an atom gains an ______ to form a more stable anion. Electron affinities are not the opposite of ionization energies, and are much ______.

H + 1312 kJ → H⁺ + e⁻ (ionization energy) – energy is required
H + e⁻ → H⁻ + 75.4 kJ (electron affinity) – energy is released

The periodic trends for electron affinity are not as regular as for ionization energy, but in general, electron affinity ______ down a group. This is because the larger the atom, the lower the electron affinity. However, as atoms get larger, the more positive nucleus ______ electron affinity. The size of the atom usually has a greater effect but there are exceptions. Across a period left to right, the electron affinity generally ______, again, with exceptions.

Periodic Trends: Electronegativity
Electronegativity is the measure of the ability of an atom to ______ electrons in a covalent bond. It is measured on a scale of 0 - 4.0 with ______, the most electronegative element, arbitrarily assigned 4.0. As we move down a group, electronegativity slightly ______ due to increased size and shielding. As we move across a period left to right, electronegativity generally ______ due to stronger nuclear charge and smaller size.

Terminology summary:
An atomic radius is the ______ of the atom.  
Ionization energy is the amount of energy needed to ______ an electron from an atom.  
Electron affinity is how much energy is given off when an electron ______ an atom. 
Electronegativity is a measure of how much each atom ______ electrons in a covalent bond.

Octet Rule
The octet rule states that chemical compounds tend to form so that each atom (by gaining, losing, or ______ electrons) has an ______ of electrons in its highest occupied energy level. For example oxygen tends to gain 2 electrons, while sodium tends to lose 1. Although there are exceptions to the octet rule, (for example hydrogen only needs to gain 1 electron to become stable; in BF₃ boron has 6 valence electrons; or __________ molecules like SF₆) it is a useful way to view covalent bonding.

Electron Dot Notation and Lewis Structures
Drawing Lewis Structures of molecules
- Count the total number of ______ electrons.
- Draw a skeleton structure using shared pairs of electrons to show ______.
- Use ______ for the central atom, otherwise use the least electronegative element.
- H is always on the ______, and in oxyacids, H is on oxygen
- Add unshared pairs of electrons so that each atom (that will accept them) is surrounded by ___ e-.
- Fill octets outside in; if there are extra electrons, give them to the _____ atom.
- Multiple pairs of electrons may be shared between 2 atoms, forming _______ and _______ bonds, to satisfy the octet rule.
- If there is more than 1 possible location for a double bond, use ___________ structures to show all possibilities; separate resonance structures with double arrows.
- Calculate the formal charge on each atom. If there is more than one way to arrange the structure, use the one with the _______ formal charge. If any formal charges are left over, include them in the Lewis structure. Central atoms from period 3 and lower can have more than 8 e\(^{-}\) (hypervalent) and can make more than 4 bonds. ex) PCl\(_5\)

\[
\text{FC} = (\text{group}) - (\text{number of } ) - (\text{number of } )
\]

Ex) :C\(=\)O :

The formal charge will _______ the charge of the species. Negative FC usually appear on the most electronegative atom, and positive FC usually appear on the _______ electronegative atom. Structures with the same sign of FC on adjacent atoms is _______.

*If and ______ if you have to __________ the octet, continue to expand the octet to reduce formal charges as much as possible.*

**Drawing Lewis Structures of Ionic Compounds**

Be sure to transfer the valence electron from the _____ to the ____. The metals will have no valence left and the non-metal will have a ______________ – always.

ex) SrF\(_2\)  

ex 2) Al\(_2\)O\(_3\)

**VSEPR Theory and Molecular Geometry** p.183

VSEPR stands for valence shell electron pair repulsion, and this theory states that the repulsion between valence electrons in molecules causes their atoms to be oriented as _______ as possible. See Handout

**Molecular Polarity** p.190

A dipole is an area where there are _______ charges separated by a short distance. A polar bond has a dipole, and it is represented by an arrow with a positive tail. If a molecule has equal dipoles that cancel each other in 3 dimensions, the molecule is _______. If the dipoles are unequal, or don't cancel, or there are lone electrons on the central atom, the molecule is _______.

\[
\text{H--Cl} \quad \text{O--C--O} \quad \text{H--N--H}
\]

molecule

\[
\text{molecule}
\]
Chem 30 Unit 2 MS2: Investigate how the properties of materials are dependent on their underlying intermolecular and intramolecular forces.

Outline
- types of bonds
- nonpolar-covalent bonds
- polar covalent bonds
- ionic bonds
- covalent bonds orb. not.
- hybridization
- sigma + pi bonds
- metallic bonds
- bond + lattice energy
- intermolecular forces

- dipole-dipole forces
- hydrogen bonding
- London dispersion forces
- surface tension, capillary
- action + viscosity
- solids: crystals, amorphous
- liquids and fluids
- gases ideal and real
- Avogadro’s Law
- Boyle’s Law

- Charles’ Law
- Gay-Lussac's Law
- Ideal Gas Law
- vapour pressure
- changes of state
- evaporation and boiling
- $\Delta H_{vap}$, $\Delta H_{fus}$
- phase diagrams

Chemical Bonds
A chemical bond is the mutual attraction between the _______ of an atom and the _______ electrons of another atom. Bonds occur because the particles have a _______ potential energy bonded to each other than they do separately.

Types of Bonds

Bonds between things vs. within things
There are forces of attraction that act _______ molecules, holding them together in a phase.
These are called ______________ forces of attraction. In order to melt or vapourize a substance, one must add enough energy to _______ these forces. The higher the melting point (or boiling point) the _______ the intermolecular forces of attraction.

The forces acting _______ a molecule holding the atoms together are called ______________ forces of attraction. These forces are much _______ than intermolecular forces and are not affected by physical changes such as adding heat. A chemical bond is an example of an intramolecular force.

Chemical bonds can be classified into 1 of 4 categories: nonpolar-covalent, polar-covalent, ionic, and metallic. We can identify the bond type based on the difference of electronegativity.

Nonpolar-covalent Bonds
Covalent means _______ valence electrons, and this occurs when two atoms have similar electronegativity values. A non-polar bond is a one where the electrons are shared _______, and no partial charges occur. These bonds occur between atoms when their electronegativity difference is between ____________ (for electronegativity values see p.151)

ex) $H_2$  \[ \text{electronegativity difference} = \frac{\text{H} - \text{H}}{2} \]  e- are shared exactly equally  (also called pure covalent at 0.0)
Polar-covalent Bonds
Polar-covalent means that there are ________ ends to the bond, and although they are sharing valence electrons, they are ______ shared equally. These bonds occur between atoms when their electronegativity difference is between ____ and ____.

ex) HCl
\[ \delta^+ \text{H} - \text{Cl}^\delta^- \]
e- are shared, but not equally; partial charges occur because of uneven electron distribution: e- are more strongly attracted to the Cl

difference =

Ionic Bonds
Ionic bonds are so ________ that electrons are actually __________ from one atom to another, creating ions. These oppositely charged particles strongly attract each other. These bonds occur between atoms when their electronegativity difference is ______ or greater. Ionic bonds form between cations and anions as a crystal lattice. This means that each cation is surrounded by anions and vice versa. This is why ionic crystals are ________, because if the lattice is shifted, strong repulsive forces cause it to shatter.

ex) NaCl
Note that NaCl is not a molecule!

\[ \text{Na}^+ - \text{Cl}^- \]

<table>
<thead>
<tr>
<th>(EN diff)</th>
<th>(0)</th>
<th>(0.3)</th>
<th>(1.7)</th>
<th>(4.0)</th>
</tr>
</thead>
</table>
\[\text{NPC PC I}\]

Covalent Bonds explained with Orbital Notation p.165
To form a bond, atoms must _______ energy; to break it, the same amount must be _______. Bond energy is the energy required to _______ a bond to form neutral atoms, and is measured in kJ/mol. The higher the bond energy, the _________ the bond. Bonding can be represented through orbital notation:

ex) fluorine can make 1 bond with itself:

\[
\begin{array}{c}
\uparrow \downarrow \\
1s & 2s & 2p
\end{array}
\]
covalent bond (overlapping orbitals)

F

\[
\begin{array}{c}
\uparrow \downarrow \\
1s & 2s & 2p
\end{array}
\]

ex 2) Nitrogen can make 3 bonds with hydrogen:

\[
\begin{array}{c}
\downarrow \downarrow \\
1s & 1s & 1s
\end{array}
\]

3 covalent bonds
Hybridization
Hybrid orbitals are orbitals of equal energy produced by ____________ 2 or more orbitals on the same atom.

<table>
<thead>
<tr>
<th>Orbitals to combine</th>
<th>Hybrid orbitals</th>
<th>Shape</th>
<th>Bond angle (°)</th>
</tr>
</thead>
<tbody>
<tr>
<td>s + p</td>
<td>linear</td>
<td></td>
<td>180</td>
</tr>
<tr>
<td>s + p + p</td>
<td>trigonal-planar</td>
<td></td>
<td>120</td>
</tr>
<tr>
<td>s + p + p + p</td>
<td>tetrahedral</td>
<td></td>
<td>109.5</td>
</tr>
<tr>
<td>s + p + p + p + d</td>
<td>trigonal bipyramidal</td>
<td></td>
<td>90, 120</td>
</tr>
<tr>
<td>s + p + p + p + d + d</td>
<td>octahedral</td>
<td></td>
<td>90</td>
</tr>
</tbody>
</table>

Hybridization explains why carbon can make 4 bonds:

\[
\text{C} \quad \uparrow \downarrow \quad \uparrow \uparrow \quad \uparrow \quad \text{Carbon should make 2 bonds...}
\]

\[
\text{C} \quad \uparrow \downarrow \quad \uparrow \uparrow \quad \uparrow \quad \uparrow \quad \text{but the s orbital and the 3 p orbitals combine}
\]

\[
\text{to make 4 sp}^3 \text{ orbitals, which are 1 part s orbital, and 3 parts p orbital.}
\]

Sigma and Pi Bonds
When hybrid orbitals make bonds, they can be labeled as sigma ( ) or pi ( ) bonds. The first bonds made are always _______ bonds. σ bonds are simply covalent bonds and are free to _______. If there are double or triple bonds, the second and third are ___ bonds. π bonds are made from __________ unhybridized p orbitals. These bonds do not ________.

Metallic Bonds
Metallic bonds form between _______ atoms, and are different from covalent or ionic bonds. Metallic bonds do not follow the octet rule, and vacant _______ orbitals overlap. This enables electrons to _______ freely throughout the entire structure. This is why metals are good _______ of heat and electricity. The bond structure is the same in all directions, which means that you can bend and reshape the material without stressing the structure, (recall that metals are _______ and _______) Metals can absorb a wide range of light frequencies, which excites their electrons. These electrons returning to their ground state gives off light, which accounts for the ________ appearance of metals.

Bond Energy and Lattice Energy
The energy required to _______ the bonds of 1 mole of molecules to form 2 moles of atoms is called its bond energy and is measured in kJ/mol. For example, H-H bond energy is 436 kJ/mol. Bond energies can vary slightly based on what other bonds the atom is making. Double and triple bonds have _______ bond energies than single bonds and are _______.

Lattice energy is like bond energy, but for _______ substances. It is the energy ________ when 1 mole of an ionic crystalline compound is formed from gaseous ions. For example, NaCl has a lattice energy of –787.5 kJ/mol. (– means energy is released)
Intermolecular Forces

Forces between molecules are generally weaker than intramolecular or ionic attractive forces. We can tell how strong intermolecular forces are by the melting point of the substance:

<table>
<thead>
<tr>
<th>substance</th>
<th>melting point (°C)</th>
<th>interparticle forces</th>
</tr>
</thead>
<tbody>
<tr>
<td>O₂</td>
<td></td>
<td>dispersion ↓</td>
</tr>
<tr>
<td>H₂S</td>
<td></td>
<td>dipole-dipole ↓</td>
</tr>
<tr>
<td>H₂O</td>
<td></td>
<td>hydrogen bonding ↓</td>
</tr>
<tr>
<td>NaCl</td>
<td></td>
<td>ionic bonding ↓</td>
</tr>
<tr>
<td>Cu</td>
<td></td>
<td>metallic bonding ↓</td>
</tr>
<tr>
<td>SiO₂</td>
<td></td>
<td>covalent bonding ↓</td>
</tr>
</tbody>
</table>

Intermolecular forces are responsible for the ________________ of matter – properties of the whole substance, not individual members.

We will study 2 types of intermolecular forces:

1. **Dipole - Dipole Forces**

Dipole-dipole forces occur between 2 _____ molecules. They are the __________ intermolecular forces but are very weak compared to ionic, metallic, or intramolecular bonds. A dipole can also induce a dipole in a nonpolar molecule by temporarily attracting some of its electrons. This attraction is __________ than a dipole-dipole force.

   \[
   \text{H}^+ - \text{Cl}^- \quad \text{---} \quad \text{H}^+ - \text{Cl}^- \quad \text{See p.191}
   \]

1b. **Hydrogen Bonding**

A particularly strong category of dipole-dipole bonding is hydrogen bonding. This only occurs where there are H-____, H-____ and H-____ bonds. The high difference in electronegativity between H and F, O, and N make very ______ molecules, and the positive H is attracted strongly to the negative end of another molecule.

   \[
   \text{H}^+ - \text{F}^- \quad \text{---} \quad \text{H}^+ - \text{F}^- \quad \text{See p.192}
   \]

This is why H₂O (has hydrogen bonding) boils at 100°C, while a similar compound H₂S, (no hydrogen bonding) boils at −60°C.

2. **London Dispersion Forces**

London dispersion forces are __________ __________ forces caused by instantaneous dipoles when electrons happen to end up on the same side of a molecule. They occur in _____ atoms and molecules, even noble gases. (also in molecules with dipole-dipole forces) London forces __________ with mass because there are more electrons, as illustrated by boiling points of:

<table>
<thead>
<tr>
<th>H₂</th>
<th>°C</th>
</tr>
</thead>
<tbody>
<tr>
<td>O₂</td>
<td>°C</td>
</tr>
<tr>
<td>Cl₂</td>
<td>°C</td>
</tr>
<tr>
<td>Br₂</td>
<td>°C</td>
</tr>
</tbody>
</table>
Properties
Intermolecular forces are responsible for properties such as surface tension, capillary action, and viscosity. Surface tension is a property of all liquids caused by the _______________ of the particles, causing them to pull together leaving the smallest possible _______________. Capillary action is caused by the attraction of the surface of a _________ to the surface of a ___________. Viscosity is a measure of how much the material ______________ flowing. All material will flow to some extent under stress. The sheer viscosity of solids is around ________orders of magnitude greater than in liquids, which is __________ than in gases.

Solids
A solid is a substance that has a definite _________ and _________. They are generally more __________ than liquids as the particles are packed even closer together, and are considered incompressible. Solids can be classified as __________ or amorphous. Crystalline solids are made of crystals and have an ordered geometric pattern. When pieces of crystal break off, they retain the geometric pattern. Two examples are salt and diamond.

There are 4 types of crystals:
____________________ - composed of positive and negative ions - are hard, brittle, good insulators and have high melting points (e.g. NaCl)
____________________ - composed of atoms covalently bonded - essentially giant molecules - are very hard and brittle, have high melting points, and are nonconductors or semiconductors (e.g. diamond)
____________________ - composed of molecules covalently bonded - low melting points, soft, good insulators (e.g. ice)
____________________ - metal atoms in ordered pattern with sea of electrons - are able to move throughout the crystal, are good conductors - varied melting points (e.g. gold)

Amorphous solids (Gk. without shape) have no geometric __________, and if pieces break off, they can have varied structures. Although their particles are arranged __________, these particles are not constantly changing position, as in a liquid. Two examples are glass and plastic.

Liquids and Fluids
A liquid is a substance that has a definite _________ and takes the _________ of its container. A fluid is a substance that can flow and therefore take the shape of its container. Gases can be considered fluid because they can flow. According to the kinetic-molecular theory, particles of solids, liquids and gases are in constant ______________.

<table>
<thead>
<tr>
<th>increasing</th>
<th>gas</th>
<th>(dipole-dipole, hydrogen bonding, and London dispersion)</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>liquid</td>
<td></td>
</tr>
<tr>
<td></td>
<td>solid</td>
<td>increasing ______________ forces</td>
</tr>
</tbody>
</table>

Liquids have higher __________ than gases and are nearly incompressible. But because their particles are in motion, liquids will __________ through other liquids. Gases will diffuse faster than liquids due to their particles moving more quickly. Solids will also diffuse, but millions of times more slowly than in liquids.
Gases: Ideal and Real
A gas is a substance that has __ definite shape and __ definite volume – expands to fill the space. Gas laws are based on some assumptions:
- gas particles are __________, have mass, are far apart, and in constant motion
- no __________ in collisions
- no forces of __________ between particles
- gas particles have ___ volume

A gas with these characteristics is called an ideal gas. However, real gases have small attractions, ___ lose energy in collisions and do actually have a small volume. The higher the pressure or the more polar the gas, the more it ______ from ideal behaviour. Most gases such as nitrogen exhibit ______ to ideal behaviour at reasonable temperatures and pressures. We will deal with all gases as ideal in this course.

Gas Laws

Avogadro’s Law: __________ volumes of gases at the same temperature and pressure contain the __________ number of particles. To standardize conditions, we commonly use Standard temperature and pressure (______) defined in your text as 0°C and 1 atm of pressure. At STP, 22.41410 L (22.4 L) of any gas contains 1 mole of gas particles. SATP (ambient) is used in thermodynamics and is 25°C and 1 atm. Note: 1 atm = 101.325 kPa = 760 mm Hg = 760 torr.

Boyle's Law: At a constant temperature, the ________ of a sample of gas increases as pressure __________ (and vice versa). P = pressure in atm k = constant (depends on the __ = k V = volume in L mass and temp. of gas)

Charles' Law: At a constant pressure, the ________ of a sample of gas increases as temperature ________ (and vice versa). V = volume in L k = constant (depends on the __ = k T = temperature in K mass and temp. of gas)

Gay-Lussac's Law: At a constant volume, the ________ of a sample of gas increases as temperature ________ (and vice versa). P = pressure in atm k = constant (depends on the __ = k T = temperature in K mass and volume of gas)

You can also use these laws to calculate changes in a sample of gas:

\[ P_1V_1 = \frac{V_1}{T_1} = \frac{P_1}{T_1} \]

Ex) A sample of argon occupies 28 mL at 1.2 atm at const. T. What is its volume at 5.4 atm?

where: \( P = \) pressure in atm \( n = \) moles of gas in mol
\( V = \) volume in L \( T = \) temperature in K
\( R = \) ideal gas constant = 0.0821 L•atm/(mol•K)

Ex) What is the volume of 5.25 mol of oxygen gas at 36.4°C and 0.750 atm of pressure?

Recall that \( n = \frac{m}{M} \) and \( D = \frac{m}{V} \). With some rearranging, we can write
\[
M = \frac{mRT}{PV} \quad \text{and} \quad D = \frac{MP}{RT}
\]

Ex 2) Find the molar mass of an unknown gas if 986 g occupies 12.7 L at 29.7°C and 60.3 atm.

Vapour Pressure
Vapour pressure is the pressure exerted by the particles in vapour phase above a liquid. The ________ the intermolecular bonds, the more easily the liquid will evaporate, and the ________ the vapour pressure. Liquids that evaporate easily (or have a higher % of liquid particles that can evaporate) are said to be ________. (e.g. ether, acetone) Note that the equilibrium vapour pressure is only dependent on temperature and is not the same as vapour pressure in an open system.

Changes of State
In this course, we will focus on the 3 states in which all matter on Earth can exist:

<table>
<thead>
<tr>
<th>least energy</th>
<th>solid</th>
<th>( \Leftrightarrow )</th>
<th>liquid</th>
<th>( \Leftrightarrow )</th>
<th>gas</th>
<th>most energy</th>
</tr>
</thead>
<tbody>
<tr>
<td>melting</td>
<td>condensation</td>
<td>sublimation</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Evaporation and Boiling
Evaporation and boiling are both forms of _________. Evaporation is when particles escape from the surface of a non-boiling liquid and enter the gas state. This happens because some of the particles on the surface of the liquid have higher than average kinetic energies, which enables them to break the intermolecular forces holding them in the liquid phase. Boiling occurs when
the liquid particles throughout the liquid have enough energy to break the intermolecular bonds → liquid and gas phases exist in equilibrium. This occurs when the liquid's equilibrium vapour pressure equals the atmospheric pressure, which is known as the ____________________.

Increasing the temperature of a liquid increases its equilibrium vapour pressure. For water:

<table>
<thead>
<tr>
<th>Temperature (°C)</th>
<th>0</th>
<th>20</th>
<th>40</th>
<th>60</th>
<th>80</th>
<th>100</th>
</tr>
</thead>
<tbody>
<tr>
<td>Vapour Press (kPa)</td>
<td>0.61</td>
<td>2.33</td>
<td>7.33</td>
<td>19.91</td>
<td>47.34</td>
<td>101.325</td>
</tr>
</tbody>
</table>

At sea level, where the atmospheric pressure is on average _______ kPa, (or 1 atm or 760 torr) water will boil at 100°C. In Banff (alt 1383 m) water will boil at 95°C and on Everest (8850 m), it will boil at 71°C. (the b.p. drops approx. _____ per 300 m) Increasing the pressure above a liquid also increases the equilibrium vapour pressure, which allows the liquid to reach a _______ temperature to boil - such as in a pressure cooker.

**Molar Heat of Vapourizaion and Fusion**

As water boils, the temperature remains ___________. The energy taken in by the water is used to break the ________________ bonds and release vapour particles. The amount of heat required to vapourize 1 mole of a liquid at its boiling point is called the molar heat of vapourization. At 1 atm, water has a molar heat of vapourization of _______kJ/mol.

The molar heat of fusion is the amount of heat required to _____ 1 mole of solid at its melting point. Freezing and condensation work the same way, with the same energies, except energy must be removed. At low temperatures and pressures, liquids cannot exist. In that case, solids can be in equilibrium with vapours, and ________________ and deposition occurs.

**Phase Diagrams**

These are pressure vs. temperature graphs showing all the phases of a substance. (See fig 12-14, p381) Note that water expands when it freezes and therefore increasing pressure lowers the melting point. This is unusual in nature. The ________________ is the temperature and pressure that a substance can exist as a solid, liquid, and gas at the same time. The ________________ is the temperature above which the substance cannot exist as a liquid, regardless of pressure. The critical pressure is the lowest pressure at the critical point.
Chem 30 Unit 3 MS3: Explore the nature and classification of organic compounds and their uses in modern materials.

Outline
- types of formulas
- isomers: structural and geometric
- naming and drawing:
  - functional groups
  - alkanes, alkenes, alkanes
- polymers: addition and condensation

Organic Chemistry
Organic chemistry is the study of compounds containing \( \text{__________} \). (except carbonates and oxides) Organic compounds are named according to a different IUPAC system. (International Union of Pure and Applied Chemistry)

Types of Formulas

<table>
<thead>
<tr>
<th>molecular</th>
<th>structural</th>
<th>condensed</th>
<th>skeletal</th>
</tr>
</thead>
<tbody>
<tr>
<td>C\textsubscript{7}H\textsubscript{16}</td>
<td>( \text{CH}_3\text{CHCH}_2\text{CH}_2\text{CH}_3 )</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Saturated Hydrocarbons - Alkanes p.634
Hydrocarbons are compounds with only hydrogen and carbon. Being saturated means that each carbon, which can make 4 bonds, has as much hydrogen on it as \( \text{__________} \). Hydrocarbons with only single bonds (saturated) are called \( \text{__________} \), and use the following prefixes to tell the \# of C.

- meth_ = hex_ =
- eth_ = hept_ =
- prop_ = oct_ =
- but_ = non_ =
- pent_ = dec_ =

methane \( \text{CH}_4 \)
ethane \( \text{C}_2\text{H}_6 \)
propane \( \text{C}_3\text{H}_8 \)
butane \( \text{C}_4\text{H}_{10} \)
pentane \( \text{C}_5\text{H}_{12} \)
hexane \( \text{C}_6\text{H}_{14} \)
heptane \( \text{C}_7\text{H}_{16} \)
octane \( \text{C}_8\text{H}_{18} \)
onane \( \text{C}_9\text{H}_{20} \)
decane \( \text{C}_{10}\text{H}_{22} \)
alkanes \( \text{C}_n\text{H}_{2n+2} \)
Adding groups

Alkyln groups are groups of atoms formed when 1 ________ is removed from an alkane. Prefixes are as before and suffixes are "-_yl". ex) methyl, butyl, propyl, etc. ex) methypropane

Alkyl halides are hydrocarbons where one or more hydrogen atom replaced by a halogen. They receive the appropriate prefix: fluoro_, chloro_, bromo_ etc. ex) chloroethane

Naming Alkanes

1. Name the longest continuous hydrocarbon chain (__________). Give it a prefix to tell how many carbons, add the suffix "ane"
2. Add the names of the alkyl groups in front of the backbone name. If there is more than 1 group, arrange them ___________. ex) ethyl methylhexane
   If there are 2 or more of the same group, use prefixes di ( ), tri ( ), and tetra ( ).
   ex) diethyloctane  Do this after they are in alphabetical order.
3. Number the backbone to give the _______ numbers, and assign position numbers to groups.
   ex) 3-ethyl, 2,4,5-trimethylhexane
4. Name halogens the same way as alkyl groups, but place them in front of them. (their higher _________ has no bearing on counting unless as a tie breaker) ex) 4-bromo-2-methylhexane

Cycloalkanes

Cycloalkanes are alkanes where the first and last carbons form a bond, creating a _______. Their empirical formula is C\textsubscript{n}H\textsubscript{2n}, because 2 hydrogen were lost to form the bond. Their name is preceded by the prefix cyclo_, ex) cyclobutane □ cyclopropane Δ

Isomers

Isomers are compounds that have the same _________ but different ___________.

Structural Isomers – (different orders)
There are 3 structural isomers of pentane: n-pentane, 2-methylbutane, and 2,2-dimethylpropane
There are 5 isomers of hexane:

Geometric Isomers – (same order, different 3-D arrangement)
Isomers differ in physical properties such as ________ points or densities. They can even have different physiologic effects – Research has suggested that ______ fats are more harmful to human health than cis fats.

Alkenes  
Unsaturated hydrocarbons have double or triple bonds, and therefore do not have as many hydrogen atoms as ______. Alkenes have at least one ______ bond, and their empirical formula is $C_nH_{2n}$. All the former naming rules still apply, with some additional considerations.
ethene  $C_2H_4$
propene  $C_3H_6$
but-1-ene  $C_4H_8$  trans-but-2-ene
pent-1-ene  $C_5H_{10}$
hex-1-ene  $C_6H_{12}$  cis-but-2-ene

Naming Alkenes  
1. Name the longest backbone that contains a double bond. Use the suffix _____, and if there is more than 1 double bond, use _adiene, _atriene, and _atetraene. ex) hexadiene
2. Number the backbone so that the double bonds have the ______ number.
3. Indicate where the _______ bonds are with a position number in front of the backbone name.

Alkynes  
Alkynes have at least one ______ bond, which is linear. As such, there is no cis or trans configurations possible. The empirical formula for alkynes is $C_nH_{2n-2}$. The naming rules for alkenes apply to alkynes, except the suffix is _yne. ex) butyne, hexa-1,4-diyne
ethyne  $C_2H_2$
propyne  $C_3H_4$
but-1-yne  $C_4H_6$
pent-1-yne  $C_5H_8$
hex-1-yne  $C_6H_{10}$

Benzene and Phenol  
Benzene is an important chemical because it is stable and is the foundation for many complex structures. Benzene is an ________ hydrocarbon, which means that it has sigma bonds and delocalized pi bonds in a ring. Its formula (cyclohexa-1,3,5-triene) would suggest 3 double bonds, but in fact all the bonds are the same length. The double bonds are shared over the whole molecule as resonance hybrid bonds. ________ is benzene with an alcohol group on carbon 1.

Three configurations are common when attaching 2 groups onto benzene: ex) with chlorine
1,2-dichlorobenzene  1,3-dichlorobenzene  1,4-dichlorobenzene
ortho-dichlorobenzene  meta-dichlorobenzene  para-dichlorobenzene
Some Other Functional Groups

1. Ethers

- longest chain is backbone, the other named as substituent with “oxy”
- numbers indicate where substituent attaches to backbone

2. Amines

- ends with “-amine” (think -anamine)
- count so N is lowest position
- root is longest chain with N
- alkyl groups on N get “N” for pos.

3. Alcohols

- ends with “-ol” (think -anol)
- count so OH group is lowest position
- if there is a higher priority functional group, name the OH group as hydroxy
4. Ketones

-ends with “-one” (think -anone)
-count so lowest number on carbonyl (C=O)

5. Aldehydes

-ends with “-al” (think -anal)
-carbonyl carbon is always carbon 1
-carbonyl carbon always has an H on end

6. Amides

-ends with “amide” (think –anamide)
-count from carbonyl carbon
-groups on N get “N” prefix
7. Esters

-ends with “oate” (think –anoate) this is the backbone name this first -use “yl”
count from carbonyl

8. Carboxylic Acids

-ends with “oic acid” (think –anoic acid)
carbonyl carbon is always number 1

<table>
<thead>
<tr>
<th>functional group</th>
<th>suffix</th>
<th>prefix</th>
<th>The suffix is used when the functional group is the highest on this list.</th>
</tr>
</thead>
<tbody>
<tr>
<td>carboxylic acids</td>
<td>-oic acid</td>
<td>carboxy-</td>
<td></td>
</tr>
<tr>
<td>esters</td>
<td>-oate</td>
<td>alkoxycarbonyl-</td>
<td></td>
</tr>
<tr>
<td>amides</td>
<td>-amide</td>
<td>amido-</td>
<td></td>
</tr>
<tr>
<td>aldehydes</td>
<td>-al</td>
<td>alkanoyl-</td>
<td></td>
</tr>
<tr>
<td>ketones</td>
<td>-one</td>
<td>oxo-</td>
<td></td>
</tr>
<tr>
<td>alcohols</td>
<td>-ol</td>
<td>hydroxy-</td>
<td></td>
</tr>
<tr>
<td>amines</td>
<td>-amine</td>
<td>amino-</td>
<td></td>
</tr>
<tr>
<td>alkenes</td>
<td>-ene</td>
<td>alkenyl-</td>
<td></td>
</tr>
<tr>
<td>(benzene)</td>
<td>benzene</td>
<td>phenyl-</td>
<td></td>
</tr>
<tr>
<td>alkynes</td>
<td>-yne</td>
<td>alkynyl-</td>
<td></td>
</tr>
<tr>
<td>alkanes</td>
<td>-ane</td>
<td>alkyl-</td>
<td></td>
</tr>
<tr>
<td>ethers</td>
<td>--</td>
<td>alkoxy-</td>
<td></td>
</tr>
</tbody>
</table>
Polymers
Polymers are large molecules made from small units called __________ joined together. Polymers are extremely important to materials science and can be found everywhere. Examples include starch, protein, __________, polystyrene, nylon, and polyester.

Addition Polymers
Addition polymers are formed when monomers with ________ bonds are joined together. For ex)

\[ n \text{CH}_2=\text{CH}_2 \rightarrow (\text{CH}_2\text{CH}_2)_n \]

Condensation Polymers
Condensation polymers are made by reactions that remove ________ to join the monomers. Examples are nylon, ________ or polyethylene terephthalate, and polyamides like Kevlar.

Organic Compound / Polymer

<table>
<thead>
<tr>
<th>Organic Compound / Polymer</th>
<th>Common Use</th>
</tr>
</thead>
<tbody>
<tr>
<td>citric acid</td>
<td>flavouring agent</td>
</tr>
<tr>
<td>dichlorodifluoromethane (Freon-12)</td>
<td>refrigerants</td>
</tr>
<tr>
<td>high density polyethylene (HDPE)</td>
<td>milk jugs, shampoo bottles</td>
</tr>
<tr>
<td>low density polyethylene (LDPE)</td>
<td>squeeze bottles, carpet, bread bags</td>
</tr>
<tr>
<td>methanal (formaldehyde)</td>
<td>biological preservation</td>
</tr>
<tr>
<td>ethane-1,2-diol (ethylene glycol)</td>
<td>antifreeze</td>
</tr>
<tr>
<td>propane-1,2,3-triol (glycerin/glycerol)</td>
<td>sweetener, thickening agent</td>
</tr>
<tr>
<td>polyamides</td>
<td>protein, nylon, Kevlar</td>
</tr>
<tr>
<td>polydichloroethylene</td>
<td>plastic wrap (Saran)</td>
</tr>
<tr>
<td>polyesters</td>
<td>clothing, plastic bottles</td>
</tr>
<tr>
<td>polyethylene terephthalate (PET/E)</td>
<td>most common thermoset plastic (is a polyester)</td>
</tr>
<tr>
<td>polypropylene (PP)</td>
<td>consumer plastics, containers, fabric, packaging</td>
</tr>
<tr>
<td>polystyrene (PS)</td>
<td>insulation, cups, packing</td>
</tr>
<tr>
<td>polytetrafluoroethylene (PFTE)</td>
<td>non-stick cookware</td>
</tr>
<tr>
<td>polyvinyl chloride (PVC)</td>
<td>plumbing pipes, siding, records</td>
</tr>
<tr>
<td>propan-2-one (acetone)</td>
<td>nail polish remover</td>
</tr>
</tbody>
</table>

Functional Group

<table>
<thead>
<tr>
<th>Functional Group</th>
<th>Examples found in:</th>
</tr>
</thead>
<tbody>
<tr>
<td>carboxylic acids</td>
<td>soap, soft drinks, vinegar, amino acids</td>
</tr>
<tr>
<td>esters</td>
<td>fruit and perfumes</td>
</tr>
<tr>
<td>amides</td>
<td>solvents, urea, pharmaceuticals</td>
</tr>
<tr>
<td>aldehydes</td>
<td>embalming, perfumes, cosmetics, dyes</td>
</tr>
<tr>
<td>ketones</td>
<td>solvents, paint and nail polish remover</td>
</tr>
<tr>
<td>alcohols</td>
<td>solvents, fuel, beer</td>
</tr>
<tr>
<td>amines</td>
<td>pharmaceuticals, dyes</td>
</tr>
<tr>
<td>alkenes</td>
<td>ripening fruit, polymers</td>
</tr>
<tr>
<td>alkynes</td>
<td>welding, polymers</td>
</tr>
<tr>
<td>alkanes</td>
<td>natural gas, fuel, wax</td>
</tr>
<tr>
<td>ethers</td>
<td>anesthetics, engine starter, refrigerant</td>
</tr>
</tbody>
</table>
Chem 30 Unit 4 EC1: Investigate the chemistry of oxidation and reduction reactions.

Outline:
- strength of oxidizing and reducing agents
- oxidation numbers
- oxidation-reduction reactions
- LEO GER
- half reactions
- redox and covalent bonds
- balancing redox reactions
- hydrogen peroxide

Outline:
- strength of oxidizing and reducing agents
- electrochemistry
- voltaic cells - wet cells, dry cells
- An Ox, Red Cat
- electrolytic cells
- electroplating
- electrode potentials
- LEO GER
- An Ox, Red Cat
- half reactions
- redox and covalent bonds
- balancing redox reactions
- hydrogen peroxide

Oxidation Numbers p.591
The distribution of electrons in a molecule can be described by oxidation numbers, which are usually _______ to the charge an atom will make. Oxidation numbers differ from charges in that they are not physically _______, but only a description of attraction for bonding electrons.
Examples: Li\(^{+1}\), oxidation states
Li\(^{1+}\), charges

Electronegativity is the ability of an atom to attract a shared pair of electrons. (See p. 151)

Rules for Assigning Oxidation Numbers
1. Uncombined elements have an Ox # of _____.
   ex) Na, C, O, P, S all have Ox # of 0
2. Elements are assigned the Ox # equal to the charge they would have as an ion. The most and least electronegative elements get their ____________.
   ex) NaOH Na is +1, O is −2 (H is +1)
3. Fluorine in a compound is always _____.
4. Oxygen is almost always ____. Exceptions H\(_2\)O\(_2\) (O = −1), OF\(_2\) (O = +2)
5. Hydrogen is ____ when bonded to a more electronegative element. When bonded to metals, hydrogen is ___.
6. Ox #s add up to 0 in a neutral compound. In a polyatomic ion, they add up to equal the _____.
7. These rules can also be applied to _______ compounds. ex) NaCl Na is +1, Cl is −1.
Use these rules to assign Ox #s to each element in a compound:
   NaF
   H\(_2\)SO\(_4\)
   PO\(_4\)\(^{3−}\)

Oxidation and Reduction p.593
Red-ox reactions involve transfer of _______. Oxidation originally described reactions of metal and oxygen to form oxides: Fe + O\(_2\) → Fe\(_2\)O\(_3\)
Reduction originally described reactions where the metal ore was purified to metal:
   Fe\(_2\)O\(_3\) → Fe + O\(_2\)
Now, we use a broader definition. An oxidation reaction is a reaction where an ion or atom ________ its oxidation state. A reduction reaction _________ the oxidation state of an element.
In this example, sodium goes from an Ox # of 0 to +1.
Ex) 2Na + Cl\(_2\) → 2NaCl Chlorine goes from 0 to −1.
Sodium is ___________ because it loses an electron.

Chlorine is reduced because it gains an electron.

Loss of Electrons is Oxidation

\[
Na \rightarrow Na^+ + e^- \\
\text{Sodium goes from 0 to +1}
\]

Gain of Electrons is reduction

\[
Cl_2 + 2e^- \rightarrow 2Cl^- \\
\text{Chlorine goes from 0 to } -1
\]

increase in Ox # = ___________
decrease in Ox # = ___________

Oxidation produces ___________ while reduction acquires them. These processes must happen together and the number of e\(^-\) produced must ______ the number acquired. The substance that is oxidized is also called a ___________ agent - it allows another substance to be reduced. The substance that is reduced is also called an ___________ agent (also oxidant or oxidizer) - it allows another substance to be oxidized.

Some Oxidizers at St. Joseph

<table>
<thead>
<tr>
<th>Oxidizer</th>
<th>Chemical Group</th>
<th>Chemical Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>NaN3</td>
<td>(NH(_4))(_2)Cr(_2)O(_7)</td>
<td></td>
</tr>
<tr>
<td>NH(_4)NO(_3)</td>
<td>(NH(_4))(_2)S(_2)O(_8)</td>
<td>peroxides</td>
</tr>
<tr>
<td>Sr(NO(_3))(_2)</td>
<td>KMnO(_4)</td>
<td>nitrates</td>
</tr>
<tr>
<td>AgNO(_3)</td>
<td>K(_2)Cr(_2)O(_7)</td>
<td>nitrites</td>
</tr>
<tr>
<td>Ba(NO(_3))(_2)</td>
<td>K(_2)CrO(_4)</td>
<td>perchlorates</td>
</tr>
<tr>
<td>KNO(_3)</td>
<td>KCIO(_3)</td>
<td>chlorates</td>
</tr>
<tr>
<td>Ni(NO(_3))(_2)</td>
<td>KIO(_3)</td>
<td>chlorites</td>
</tr>
<tr>
<td>Mg(NO(_3))(_2)</td>
<td>Na(_3)BO(_3)</td>
<td>hypochlorites</td>
</tr>
<tr>
<td>Pb(NO(_3))(_2)</td>
<td>MnO(_2)</td>
<td>dichromates</td>
</tr>
<tr>
<td>LiNO(_3)</td>
<td>H(_2)O(_2)</td>
<td>permanganates</td>
</tr>
<tr>
<td>Cu(NO(_3))(_2)</td>
<td>Ca(NO(_3))(_2)</td>
<td>persulfates</td>
</tr>
</tbody>
</table>

Ex 2) Silver nitrate acts on copper to form copper(II) nitrate and silver. Identify the species that is oxidized and reduced.

\[
\text{Cu} + \text{AgNO}_3 \rightarrow \text{Cu(NO}_3)_2 + \text{Ag} \\
\text{Ag is } ___________ \text{ (ox # decreases)} \\
\text{Cu is } ___________ \text{ (ox # increases)}
\]

Ex 3) Identify the reducing and oxidizing agents.

\[
\text{Zn} + \text{HNO}_3 \rightarrow \text{Zn(NO}_3)_2 + \text{NO}_2 + \text{H}_2\text{O} \text{ (unbalanced)}
\]
<table>
<thead>
<tr>
<th>Element</th>
<th>Ox # change</th>
<th>e⁻ change</th>
<th>process</th>
<th>agents</th>
</tr>
</thead>
<tbody>
<tr>
<td>Zn</td>
<td>loss of 2</td>
<td></td>
<td></td>
<td>Zn is a _________ agent</td>
</tr>
<tr>
<td>N</td>
<td>gain of 1</td>
<td></td>
<td></td>
<td>HNO₃ is an _________ agent</td>
</tr>
</tbody>
</table>

(note: Nitrates are oxidizers, Zn(NO₃)₂ too is a strong oxidizer)

Pennies will react with nitric acid to produce NO in the following redox reaction:

$$3Cu + 2NO_3^- + 8H^+ \rightarrow 3Cu^{2+} + 2NO + 4H_2O$$ (you write in the Ox #s)

$$3Cu \rightarrow 3Cu^{2+} + 6e^-$$

oxidation half reaction

$$2NO_3^- + 6e^- + 8H^+ \rightarrow 2NO + 4H_2O$$

reduction half reaction

(nitrate is reduced to NO)

2NO + O₂ → 2NO₂
dark brown gas in air (smog) light brown in water nitric acid (acid rain)

Note that the following is not a redox reaction because there is no change to oxidation numbers:

$$NaCl + AgNO_3 \rightarrow NaNO_3 + AgCl$$

**Redox and Covalent Bonds**

Oxidation-reduction occurs in covalently bonded molecules as well.

$$H_2 + Cl_2 \rightarrow 2HCl$$

Even though there is no loss or gain of e⁻, they are not shared. Here, +1 describes a partial transfer of an e⁻

In other words, the shared pair of e⁻s hangs around Cl more than the H.

**Balancing Redox Reactions**

There are many methods developed to balance redox reactions. Often they are difficult to balance by inspection because you must balance both the number of atoms and oxidation states.

**Oxidation Number Method:**

This method is basically inspection, but we get a starting position from the Ox #s.
Ex 2) \( \text{NH}_3 + \text{O}_2 \rightarrow \text{NO}_2 + \text{H}_2\text{O} \)

Ex 3) \( \text{Zn} + \text{HNO}_3 \rightarrow \text{Zn(NO}_3)_2 + \text{NO}_2 + \text{H}_2\text{O} \)

**Balancing Redox Reactions in Acidic Conditions**
Without going into the half reaction method, here is a short-cut method to balance these:
1) balance as before
2) If the H and O don't balance, balance the O first by adding \( \text{H}_2\text{O} \) to one side
3) Add \( \text{H}^+ \) to the needed side to balance H

Ex) Balance \( \text{MnO}_4^- + \text{Fe}^{2+} \rightarrow \text{Mn}^{2+} + \text{Fe}^{3+} \), which takes place in acidic conditions.
Balancing Redox Reactions in Basic Conditions

Balance the equation as if it were in acidic condition. If there is H⁺ left over, then add OH⁻ to both sides to cancel the H⁺. Finally, combine the H⁺ and OH⁻ to form H₂O.

Ex) Balance Cr(OH)₃ + BrO₃⁻ → CrO₄²⁻ + Br⁻ in basic conditions.

Strength of Oxidizing and Reducing Agents

Recall that:

<table>
<thead>
<tr>
<th>substance</th>
<th>Ox #</th>
<th>electrons</th>
<th>category</th>
<th>example</th>
</tr>
</thead>
<tbody>
<tr>
<td>oxidized</td>
<td>loss of e⁻</td>
<td>agent</td>
<td>Li → Li⁺</td>
<td></td>
</tr>
<tr>
<td>reduced</td>
<td>gain of e⁻</td>
<td>agent</td>
<td>Ag⁺ → Ag</td>
<td></td>
</tr>
</tbody>
</table>

The strength of oxidizing and reducing agents can be compared based on their abilities to give and take __________. The easier a substance gives up electrons, the stronger a __________ it is. The more strongly a substance attracts electrons, the stronger an __________ it is. These properties are summarized by the activity series and are related to electronegativity. (See table 19-3, p.603) Any reducing agent can be oxidized by the oxidizers __________ it, and vice versa. The farther apart the two, the more likely a redox reaction will take place.

Hydrogen Peroxide

2HOOH → 2HOH + O₂

The peroxide ion O₂²⁻ can easily be both oxidized and reduced. This is autooxidation -self red-ox.

Electrochemistry

Electrochemistry involves changes in chemical and electrical energy. We will study two electrochemical cells: __________ cells and __________ cells.

Voltaic Cells

A voltaic electrochemical cell (also called galvanic cell) __________ converts chemically stored potential energy into moving electrons through a redox reaction. (Ex - battery) If this reaction occurs in contact, such as in a beaker, there is an __________ in temperature. However, if the electrodes are separated by a conductor, useful electrical energy can be produced. An
An electrochemical cell is made of two electrodes; one is oxidized (the [underline]_________), the other is reduced (the [underline]_________). The classical example of this is a zinc-copper cell. Because Zn loses its electrons more easily than Cu, Zn is oxidized, while Cu is reduced.

\[
\begin{align*}
\text{Zn} & \rightarrow \text{Zn}^{2+} + 2e^- \\
\text{Cu}^{2+} + 2e^- & \rightarrow \text{Cu}
\end{align*}
\]

Oxidation Reduction
Anode (\(-\)) Cathode (+)

A voltaic wet cell:

Notice that the electrodes are in an electrolytic solution with the same metal as an ion. In order for the \(e^-\) to flow, we must complete the circuit with a [underline]_________ - a path for ions to migrate in either direction, but keeping the metal atoms from doing the same.

Zn loses \(e^-\)s and forms \(\text{Zn}^{2+}\) in solution
\(\text{Cu}^{2+}\) gains \(e^-\)s from the copper and grows
Zn [underline]_________, while copper grows [underline]_________

With excess \(\text{Zn}^{2+}\), the reaction would stop but the salt bridge allows more \(\text{NO}_3^-\) to cross over, keeping it electrically [underline]_________.

\(\text{K}^+\) performs the same function on the other side.

Electrolytic Cells

Electrolytic cells are cells where electricity is used to [underline]________ a spontaneous redox process. Two examples of electrolysis are recharging batteries and electroplating. Because electrolysis is [underline]spontaneous, it requires an external voltage source to reverse the reaction. A good example of this is the car battery. This battery is made of 6 cells of \(\text{Pb}\) and \(\text{PbO}_2\). Under oxidation of the \(\text{Pb}\) and reduction of \(\text{PbO}_2\) in sulfuric acid, both electrodes are coated with \(\text{PbSO}_4\) as the battery runs down. In this case, the battery acts as a voltaic cell, as it produces 12V. However, the car's alternator (and a rectifier) produce slightly more than 12V in the opposite direction. This
reverses the redox reaction, and converts the PbSO₄ back to Pb and PbO₂. Here, the battery is acting as an ____________ cell.

Electroplating p.611
In a voltaic cell, the anode ___________, while the cathode is plated and grows larger. If a DC voltage is applied (negative terminal connected to the former anode) the process is ___________. Electroplating uses electricity to change the ___________ to a ___________ which now accepts ions from the solution causing metal deposition on its surface.

Electrode Potentials p.613
Metals tend to lose e⁻'s in reactions and are therefore ___________. Some metals lose e⁻'s more easily than others as was seen in the activity series. The tendency of any half reaction to occur as a reduction reaction is called the ___________. We can also quantify this characteristic by measuring the potential difference (emf or E) in volts produced between any two electrodes. This is called an ___________. To standardize these half reaction voltages, we set the half reaction of ___________ to zero volts:

\[ 2H^+ + 2e^- \rightarrow H_2 \quad E^o = 0.00 \text{ V} \quad (\text{recall: } ^o = \text{standard states: } 25^\circ \text{C}, 1 \text{ atm, } 1 \text{M}) \]

A table of these reduction half reactions shows the standard electrode potentials relative to the standard hydrogen electrode (SHE) See table 19-3, p.615 for values. Of the substances listed, Li most easily ___________ e⁻'s, while F₂ most easily ___________ them. Note: they are all written as reductions. Sample reduction half reactions:

\[ \text{Cu}^{2+} + 2e^- \rightarrow \text{Cu} \quad E^o = +0.34 \text{ V} \quad \text{A spontaneous cell made from these metals would have copper act as the ___________ and zinc as the ___________}. \]

\[ \text{Zn}^{2+} + 2e^- \rightarrow \text{Zn} \quad E^o = -0.76 \text{ V} \quad \text{Cathode is the more (+) and anode the more (−)}. \]

To tell if a cell will produce useful electricity, we can calculate the potential difference for the cell at standard states (E°_cell)

If the E°_cell is (+), the reaction is ___________.
If the E°_cell is (−), the reaction is ___________.
Note: this is opposite to ΔH and ΔG, but same as ΔS
Ex 1) Determine the voltage of a cell made from Mg | Mg$^{2+}$ and Cu | Cu$^{2+}$. Will this process be spontaneous?

Note that the reaction will always be spontaneous in one direction.
Cu as cathode:
Cu$^{2+}$ + Mg → Cu + Mg$^{2+}$  \[ E^\circ = +2.71 \text{ V} \ (\underline{\text{_____________}}) \]

Mg as cathode:
Mg$^{2+}$ + Cu → Mg + Cu$^{2+}$  \[ E^\circ = -2.71 \text{ V} \ (\underline{\text{_____________}}) \] -could occur as electrolysis

Ex 2) Determine the voltage of the following reaction and state whether it is spontaneous.
2Al$^{3+}$ + 3Sn → 2Al + 3Sn$^{2+}$

\[
\begin{align*}
\frac{\text{2Al}^{3+} + 3\text{Sn}}{\text{2Al} + 3\text{Sn}^{2+}} & \quad E^\circ = -1.66 \text{ V} \\
\text{the more neg will be the anode in the spontaneous direction, so no – this is not spontaneous}
\end{align*}
\]

Also \[ E^\circ = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = \]

This reaction would be \underline{_____________}. However, if we applied at least 1.52 V with the negative terminal connected to Al, this reaction would occur as \underline{_____________}.

When the cell is not under standard state conditions, we must use the \underline{_____________} equation:

\[
E_{\text{cell}} = E^\circ - \frac{RT}{nF} \ln Q
\]

\[ E^\circ = \text{standard state potential} \quad R = \text{gas constant} \quad T = \text{temp in K} \]

\[ n = \text{number of e}^- \quad F = \text{Faraday constant} \quad Q = [\text{ion out}]/[\text{ion in}] \]
Chem 30 Unit 5 EQ1: Consider the characteristics and applications of equilibrium systems in chemical reactions.

Outline:
- equilibrium
- calculating $K_{eq}$
- Le Châtelier's Principle
- equilibrium shift: press. conc. temp.
- favoured reactions
- factors affecting $K_{eq}$
- the equilibrium constant $K_{eq}$
- ICE method

Equilibrium
An equilibrium is a dynamic condition in which the forward rate of a process ______ the reverse rate in a closed system. Imagine two closed containers of gas are connected and the valve separating them is opened. Each gas would flow into the other until the _____ of each gas entering and leaving becomes equal. This is equilibrium. When you put the lid on a bottle of water, the particles evaporating and condensing will reach equilibrium. (fig 12-10, p.373)

$$\text{liquid} + \text{heat} \rightleftharpoons \text{vapour}$$

**Equilibrium is shown by a __________ arrow which means that the process occurs in both directions.**

H$_2$(g) + I$_2$(g) $\rightleftharpoons$ 2HI(g)

Favoured Reactions
Recall that chemical equilibrium involves a __________ reaction in which the forward rate equals the reverse rate, and therefore the concentrations of reactants and products remain constant. So far, we have assumed that neither forward nor reverse reactions were favoured, and the concentrations of reactants and products are roughly equal at equilibrium.

If however, the forward reaction is __________, the forward reaction is nearly __________ before the reverse reaction establishes equilibrium. In this case, there is a higher concentration of __________ than reactants at equilibrium.

$$\text{HBr(aq)} + \text{H}_2\text{O(l)} \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{Br}^-(aq)$$

___________ favoured

If the reverse reaction is favoured, the forward reaction is just __________ when the reverse reaction establishes equilibrium. Here, there is a higher concentration of __________ than products at equilibrium.

$$\text{H}_2\text{CO}_3(aq) + \text{H}_2\text{O(l)} \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{HCO}_3^-(aq)$$

___________ favoured

Le Châtelier's Principle
Because equilibria are __________, any change to the reactants or products will change the equilibrium. Le Châtelier's principle states that when an equilibrium is stressed, it will ______ to relieve the stress. For example: if this system is stressed by adding more heat, the equilibrium will shift to the _____, which means that it will produce more vapour until it reaches a new ______________.

$$\text{liquid} + \text{heat} \rightleftharpoons \text{vapour}$$

Stresses include changes in __________, pressure, temperature, volume, etc. (table 12-3, p.375)

Equilibrium Shift
Recall that Le Chatelier’s principle states that an equilibrium will shift to relieve a stress.
Pressure: Increasing the pressure will shift the eq. toward the side which has fewer moles of gas particles. If they are equal, or there are no gases, there is no change. 

\[ \text{N}_2(g) + 3\text{H}_2(g) \rightleftharpoons 2\text{NH}_3(g) \quad \text{(Haber process)} \]

increase pressure, shift to right

Note: adding a non-reacting gas to increase the pressure does not affect the eq.

Concentration: Increasing the concentration of one reactant or product will shift the eq. toward the _________.

\[ \text{HBr}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{Br}^- (aq) \]

increase [HBr], shift to ________

increase [H\text{2O}], ________

Temperature: Increasing the temperature shifts the eq. ________ from the energy term.

\[ \text{N}_2(g) + 3\text{H}_2(g) \rightleftharpoons 2\text{NH}_3(g) + 92 \text{kJ} \]

increase temp, shift to ________

\[ \text{CaCO}_3(s) + 556 \text{kJ} \rightleftharpoons \text{CaO}(s) + \text{CO}_2(g) \]

increase temp, shift to ________

Note: \( K_{eq} = [\text{CO}_2(g)] \)

The Equilibrium Constant

The equilibrium constant (K or \( K_{eq} \)) describes the extent to which the reaction has been ________ by the time it reaches equilibrium. In a reaction:

\[ \text{aA} + \text{bB} \rightleftharpoons \text{cC} + \text{dD} \]

\[ K_{eq} = \left( \frac{[\text{cC}]^c[\text{dD}]^d}{[\text{aA}]^a[\text{bB}]^b} \right) \]

\[ K_{eq} = \frac{[\text{products}]}{[\text{reactants}]} \]

In general

If \( K_{eq} = >1 \) ________ favoured (more products around at eq.)

\( K_{eq} = 1 \) equal conc. of reactants and products

\( K_{eq} = < 1 \) ________ favoured (more reactants around at eq.)

NOTE: Changing the pressure, concentration, or adding a catalyst ________ the \( K_{eq} \). For example, more product may be formed (as in a shift to the right), but the ratio of their concentrations remains the same.

Changing the temperature ________ change the \( K_{eq} \). For an exothermic reaction, increasing the temp ________ the \( K_{eq} \); For an endothermic reaction, increasing the temp ________ the \( K_{eq} \).

Reactions run to completion when a product in not available for a reverse reaction. Examples of this include producing a gas that escapes, a precipitate that is insoluble, or a product that does not ionize.

The value of \( K_{eq} \) is found experimentally and is dependent on ________. If gases are used, pressure takes the place of concentration, however your text uses concentration. The conc. of pure solids and liquids essentially remain ________ and are left out of equilibrium expressions.
Ex 1) Calculate the equilibrium constant for the reaction at constant temperature -
\[ \text{N}_2 + \text{O}_2 \rightleftharpoons \text{NO} \]
if the concentrations at equilibrium are: \([\text{N}_2] = 6.4 \times 10^{-3} \text{ mol/L}, [\text{O}_2] = 1.7 \times 10^{-3} \text{ mol/L}, [\text{NO}] = 1.1 \times 10^{-5} \text{ mol/L} \).

Ex 2) Given: \( \text{H}_2(g) + \text{Br}_2(g) \rightleftharpoons 2\text{HBr}(g) \), \( K_{eq} = 3.5 \times 10^4 \), and constant temp.
a) Calculate the concentration of \( \text{H}_2 \) if \([\text{HBr}] = 9.8 \times 10^{-2} \text{ M}, \text{and} [\text{Br}_2] = 4.3 \times 10^{-3} \text{ M} \).

b) Calculate the concentration of \( \text{HBr} \) if \([\text{H}_2] = 0.0356 \text{ M} \text{ and} [\text{Br}_2] = 0.0298 \text{ M} \)

The ICE method
When reactants are converted into products, their concentrations _________ and the concentrations of the products _________ according to molar ratios. When calculating \( K_{eq} \), we must use equilibrium concentrations, not initial concentrations. To find equilibrium concentrations, it is sometimes useful to use the ICE method.

Ex 1) Given a 1.0 M HCl solution produces \([\text{H}_3\text{O}^+] \) of 0.9 M at equilibrium, what is the equilibrium concentration of HCl?

1b) Note that you can now find the \( K_{eq} \) because the equilibrium concentrations of \( \text{H}_3\text{O}^+ \) and \( \text{Cl}^- \) will be the same.
Ex 2) Given: \(2\text{Al}^{3+}(aq) + 3\text{S}^{2-}(aq) \rightleftharpoons \text{Al}_2\text{S}_3(aq)\)

a) If the concentration of \(\text{S}^{2-}\) changes from 0.45 M to 0.30 M at equilibrium, what is the equilibrium concentration of \(\text{Al}_2\text{S}_3\)?

\[
\begin{align*}
\text{I} & \quad \text{C} \\
\text{E} & \\
\end{align*}
\]

b) If the equilibrium concentration of \(\text{Al}^{3+}\) is 0.70 M, what was the initial concentration of \(\text{Al}^{3+}\)?

c) Find the \(K_{eq}\).

More Practice: Consider the following equilibrium

\[
4\text{NH}_3(l) + 5\text{O}_2(g) \rightleftharpoons 4\text{NO}(g) + 6\text{H}_2\text{O}(l)
\]

a) If the reaction proceeds from a mixture and the concentration of \(\text{NO}\) increases from 1.02 M to 2.29 M at equilibrium, find the initial concentration of oxygen if its equilibrium concentration is 2.49 M.

b) If the initial concentrations of ammonia and water are both 1.00 M, calculate the \(K_{eq}\).
Chem 30 Unit 6 EQ2: Analyze aqueous solution equilibria including solubility-product constants.

Outline
- Solubility
- common ion effect
- precipitation reactions
- calculating $K_{sp}$
- electrolytes/nonelectrolytes
- using $K_{sp}$ to find solubility
- strong/weak electrolytes
- predicting precipitates
- dissociation/ionization

Solubility
Solubility describes the amount of _______ per amount of _______ required to produce a saturated solution at a given temperature. It is usually measured in g of solute per g solvent. A saturated solution has the _______ amount of solute dissolved in that solvent at that temperature. For example, if you add sugar to water until some remains undissolved at the bottom, it is saturated. (sucrose has a solubility of 203.9 g/100 g water at 20°C) The solution has reached solution equilibrium, where dissolution and crystallization occur at the same rate. The rate at which things dissolve can be affected by changes in ____________, ____________, and ______________.

A supersaturated solution contains ____ solute than a similar saturated solution under the same conditions. This is usually done by heating the solution, adding the solute and cooling it slowly.

Electrolytes and Nonelectrolytes
An electrolyte is a substance that when dissolved in water, forms ions which ___________ an electric current (e.g. NaCl). Ionic compounds, acids and bases are electrolytes. Some polar molecules will also form ions in solution. A nonelectrolyte is a substance that when dissolved in water __________________________ and will not conduct an electric current (e.g. sucrose).

Strong and Weak Electrolytes
A strong electrolyte is a compound that breaks up (dissociates or ionizes) ____________ or almost entirely and exists as ions in solution. For example, HCl is 100 % ionized in water. Note that this is independent of solubility, and even slightly soluble compounds can be strong electrolytes (what does dissolve is in ion form)

$$\text{HCl (aq)} + \text{H}_2\text{O} (l) \rightarrow \text{H}_3\text{O}^+ (aq) + \text{Cl}^- (aq) \quad \text{strong electrolyte} \quad - \quad \text{______________}$$

A weak electrolyte has only a small amount of the dissolved compound existing as ions.

$$\text{HCN (aq)} + \text{H}_2\text{O} (l) \rightleftharpoons \text{H}_3\text{O}^+ (aq) + \text{CN}^- (aq) \quad \text{weak electrolyte} \quad - \quad \text{______________}$$

Dissociation and Ionization
When ionic compounds dissolve, they dissociate into their _______. (assume 100 % dissociation)

$$\text{CaCl}_2 (s) \quad \rightarrow \quad \text{H}_2\text{O}$$

Some examples of dissociation eqns

$$\text{Ni}_2(\text{SO}_4)_3 (s) \quad \rightarrow$$
Ionization is different in that ions are produced from substances. Ionization occurs when the intramolecular force in the solute is weaker than the attraction to the solvent. The attraction of the HCl molecule to water is stronger than the H-Cl bond.

\[
\text{HCl (aq) + H}_2\text{O (l)} \rightarrow
\]

**Precipitation Reactions** p.427
When a product of a reaction is insoluble, it precipitates out to form a solid. To predict whether a precipitate will form, use a solubility chart. (such as table 14-1, p.427)

\[
\text{Zn(NO}_3\text{)}_2(\text{aq}) + (\text{NH}_4)_2\text{S (aq)} \rightarrow
\]

\[
3\text{CuCl}_2(\text{aq}) + 2\text{Na}_3\text{PO}_4(\text{aq}) \rightarrow
\]

The overall ionic equation is

The spectator ions are not involved in the precipitation:

The net ionic equation includes only the ions that make the precipitate:

**Common Ion Effect** p.567
Introducing an ion that is common to 2 solutes will produce a precipitate or reduce ionization. For example, placing HCl into a saturated solution of NaCl will precipitate out NaCl.

\[
\text{HCl(aq) + H}_2\text{O(l)} \rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \quad \text{Adding HCl to NaCl will}
\]

\[
\text{NaCl(s)} \rightleftharpoons \text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \quad \text{Adding NaCl to HCl will}
\]

\[
\text{Note: Higher [Cl}^{-}\text{] but lower [Na}^{+}\text{] } \rightarrow \text{ same } K_{eq}
\]

CH\text{}_3\text{COOH(aq) }{+}\text{H}_2\text{O(l)} \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{CH}_3\text{COO}^-(\text{aq})
\]

Adding NaCH\text{}_3\text{COO will reduce the ionization of CH}_3\text{COOH and lower [H}_3\text{O}^+\text{] and raise pH.}

**The Solubility Product Constant** \(K_{sp}\) p.577
\(K_{sp}\) is equilibrium constant of a saturated ionic solution. In the equilibrium:

\[
\text{AgCl (s)} \rightleftharpoons \text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \quad K_{sp} = \text{ the amount of solid remains constant}
\]

\[
K_{sp} =
\]

Don't forget that the concentrations are raised to the power of their coefficient:

\[
\text{MgF}_2(\text{s}) \rightleftharpoons \text{Mg}^{2+}(\text{aq}) + 2\text{F}^-(\text{aq}) \quad K_{sp} =
\]

39
Ex 1) Calculate the $K_{sp}$ of CuCl at 25°C if its solubility is 0.0108 g/100. g H$_2$O.

Ex 2) Calculate the $K_{sp}$ for PbCl$_2$ if it has a solubility of 1.0 g/100.0 g at a particular temperature.

Using $K_{sp}$ to Calculate Solubilities

Recall that a large $K$ means that the concentrations of the products are ______ than the reactants. (or the reaction goes to completion) A small $K$ means that the reaction does not produce much product, and most remains unreacted. The $K_{sp}$ works the same way. A large $K_{sp}$ means that there are many __________ in solution (i.e. it is ________). A low $K_{sp}$ means that there are very ______ ions in solution (i.e. it is not very soluble or ________). The higher the $K_{sp}$, the ______ the solubility. For example, PbCl$_2$ has a $K_{sp}$ of $1.6 \times 10^{-5}$, (large) and is soluble at room temp. In contrast, HgS has a $K_{sp}$ of $1.6 \times 10^{-52}$, (small) and is insoluble at room temp.

Given a $K_{sp}$, you can find the solubility of a sparingly soluble salt.
Ex) Calculate the solubility of BaCO$_3$ if its $K_{sp}$ is $5.1 \times 10^{-9}$.

Predicting Precipitates

If the product of the [ions] for a slightly soluble salt exceeds the $K_{sp}$, a ______________ forms.
Ex) Will a precipitate form if 20.0 mL of 0.010 M BaCl$_2$ is mixed with 20.0 mL of 0.0050 M Na$_2$SO$_4$ if the $K_{sp}$ for BaSO$_4$ is $1.1 \times 10^{-10}$? NaCl is soluble, but BaSO$_4$ is sparingly
UNIT 7

Chem 30 Unit 7 EQ3: Observe and analyze phenomena related to acid-base reactions.

Outline
-Acid names
-Arrhenius acids and bases
-strength of acids and bases
-Brønsted-Lowry acids and bases
-polyprotic acids
-Lewis acids and bases
-conjugate acids and bases
-amphoteric/amphiprotic
-neutralization reactions

-ionization constant of water:
-\( K_w = [H_3O^+] \cdot [OH^-] = 1.0 \times 10^{-14} \text{ M}^2 \)
-calculating concentrations
-pH, pOH
-titration
-calculations with pH and titration
-K_a and pK_a
-K_b and pK_b
-calculations of K_a, pK_a, K_b, and pK_b

Acid Nomenclature p.454
Binary acids contain hydrogen and one of the more electronegative elements. Oxyacids contain hydrogen, oxygen and a third element, usually a nonmetal. Some common acids:

<table>
<thead>
<tr>
<th>Binary acids</th>
<th>Oxyacids</th>
</tr>
</thead>
<tbody>
<tr>
<td>hydrofluoric acid</td>
<td>acetic acid</td>
</tr>
<tr>
<td>hydrochloric acid</td>
<td>carbonic acid</td>
</tr>
<tr>
<td>hydrobromic acid</td>
<td>nitric acid</td>
</tr>
<tr>
<td>hydriodic acid</td>
<td>phosphoric acid</td>
</tr>
<tr>
<td>hydrosulfuric acid</td>
<td>sulfuric acid</td>
</tr>
</tbody>
</table>

The prefixes and suffixes refer to the number of oxygen atoms on the central atom:
- hypochlorous acid
- chlorous acid
- chloric acid
- perchloric acid

-hypo____ous means "2 less oxygen"
-____ous means "1 less oxygen"
-____ic is the most common form or first discovered
-per____ic means "1 more oxygen"

Arrhenius Acids and Bases p.459
According to Arrhenius, an acid is a compound that will increase the ____ in solution; A base is a compound that will increase the ____. Compounds will ionize or dissociate into these ions, or produce them from water. For example, HCl will ionize into H^+ and Cl^- (acid). Actually, the polar water molecule attracts the H^+ ion and forms H_3O^+, the hydronium ion.

Strong acids and bases ionize or dissociate _______ and are also strong electrolytes. The more polar the bond with hydrogen, the easier it is to break and form ions \( \rightarrow \) strong acid. Weak acids and bases only ionize or dissociate _________, and are weak electrolytes.

\[
\begin{align*}
\text{HF (g) + H}_2\text{O (l)} & \rightleftharpoons \text{HF}_3\text{O}^+ \quad \text{weak acid (do not confuse w/ corrosiveness)} \\
\text{NH}_3 (g) + \text{H}_2\text{O (l)} & \rightleftharpoons \text{NH}_4\text{OH} \quad \text{weak base}
\end{align*}
\]
The ___________ of the compound also affects the strength of acids and particularly bases. For example, if a metal hydroxide is insoluble, it will not produce many OH− ions.

Limitations of Arrhenius definition:
-it predicted that compounds that did not produce H+ or OH− ions would be __________
-it did not account for acids or bases not in __________ solution

Brønsted-Lowry Acids and Bases p.464
In 1923 Brønsted and Lowry independently expanded Arrhenius' definition. A Brønsted-Lowry acid is a molecule or ion that is a ______________ (H+ is a proton). As well, a base is a ______

\[
\text{HCl} + \text{NH}_3 \rightarrow \text{NH}_4^+ + \text{Cl}^- \quad \text{HCl is the ________, NH}_3 \text{ is the ____________}
\]

Arrhenius hydroxide bases such as NaOH are ______ Brønsted-Lowry bases, as they do not accept a proton. The OH− ion would be the base.

- HCl __________ acid (can donate 1 proton/molecule)
- H₂SO₄ __________ acid \ (also called polyprotic)
- H₃PO₄ __________ acid /

ionization series:
(weak acid) \[ \text{H}_3\text{PO}_4 + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \quad \text{greatest conc.} \quad \text{note: protons are} \]
(weaker acid) \[ \text{H}_2\text{PO}_4^- + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \quad \text{lower conc.} \text{donated one at a time} \]
(weaker still) \[ \text{HPO}_4^{2-} + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \quad \text{lowest conc.} \]

Limitations of Brønsted-Lowry definition:
-compounds could only be acids if they contained __________

Lewis Acids and Bases p.467
In 1923 also, a yet broader definition of acids and bases was proposed by Gilbert Lewis. A Lewis acid is an atom, molecule, or ion that accepts an __________ to form a covalent bond.

(Acid Accepts) A Lewis base is an electron pair __________.

A co-ordinate covalent bond is where both e− come from the same atom. Note that H+ acts as a Lewis __________, because it can accept an e− pair. We will use the Brønsted-Lowry definition of acids and bases unless otherwise stated.

Summary: ___________________________ Acid ___________________________ Base ________

Arrhenius
Brønsted-Lowry
Lewis

Table 15-5 p.468
Conjugate Acids and Bases

When an acid donates a proton (H\(^+\)), it creates a ________ base, which is what is left over after losing a proton.

HF\((g)\) + H\(_2\)O \((l)\) ⇌ H\(_3\)O\(^+\)(aq) + F\(^-\)(aq)  

A conjugate base can accept H\(^+\) to reform that acid

NH\(_3\)\((g)\) + H\(_2\)O \((l)\) ⇌ NH\(_4\)\(^+\)(aq) + OH\(^-\)(aq)  

A conjugate acid can release H\(^+\) to reform that base

The stronger the acid, the more it ionizes and the ______ likely the reverse reaction will occur. Therefore, the stronger the acid or base, the ______ the conjugate base or acid will be. See table 15-6 on p.471.

Amphoteric

In the above examples, water can act as either an acid or a base, and is called ______________. (water is amphoteric)

acts as acid \[\text{H}_2\text{O}\] acts as base
when reacted with when reacted with

This is easily confused with amphiprotic, which means that it can donate or accept ____. All amphiprotic substances are amphoteric, but there are substances that can act as either acid or base that do not have hydrogen, and can't be called amphiprotic. (fr Gk, amphi - both)

Neutralization

In aqueous solutions, neutralization involves the reaction of hydronium and hydroxide ions to form water and a salt, which is an ionic compound made from the cation from a base and the anion from an acid.

HCl \((aq)\) + NaOH \((aq)\) →

overall ionic eqn

net ionic eqn

Ionization of Water

Pure water will actually ionize on its own. This makes water a weak electrolyte.

H\(_2\)O \((l)\) + H\(_2\)O \((l)\) ⇌

But the concentration of its ions [H\(_3\)O\(^+\)], [OH\(^-\)] is extremely small. [X] means the concentration of X in mol/L. At room temperature, both [H\(_3\)O\(^+\)] and [OH\(^-\)] = __________ M. As the [H\(_3\)O\(^+\)] increases, the [OH\(^-\)] ________, and their product is __________. This is known as the ionization constant of water. (K\(_w\))

K\(_w\) =

(at 25°C)
Calculating Concentrations

Because \([H_3O^+] = [OH^-]\) for water, it is neutral. However, if a solution has a greater \([H_3O^+]\) than \([OH^-]\), it is _______. For bases, \([OH^-] > [H_3O^+]\).

Ex 1) Calculate the \([H_3O^+]\) and \([OH^-]\) for a 5.0 \(\times 10^{-2}\) M solution of HI.

Ex 2) What is the \([H_3O^+]\) for a 3.0 M solution of \(H_2SO_4\)? (we will assume 100% ionization)

\[
\text{pH} = \quad \text{also:} \quad \text{pOH} =
\]

The pH scale generally ranges from _______, with ____ being neutral. pH below 7 is _______, and above is _______. Note that pH + pOH = _____. See table 16-3 p.486 for examples.

Ex 1) calculate the pH of a 2.5 \(\times 10^{-2}\) M solution of HCl.

Significant figures with pH are a little different. The number of digits after the decimal determines the number of sig figs in a pH value.

ex 2) Calculate the \([H_3O^+]\) of a solution with a pOH of 8.0.

Titration

So far, we have only been dealing with strong acids and bases (complete ionization/dissociation). Weak acids and bases _______ completely break into their ions, and can't be calculated this way. An approximation of the pH can be found experimentally using and indicator or more accurately by titration. Titration is the process of mixing an acid with a base to find the ______________ - when the concentration of \([H_3O^+]\) = \([OH^-]\). If you know the concentration and volume of one reactant (the standard solution), you can determine the __________ of the other, knowing the volume required to "neutralize" it. For a strong acid (like HCl) and a strong base (like NaOH), the equivalence point occurs around pH = ___. However, when you titrate a weak acid
(like acetic acid) with a strong base like (NaOH), the equivalence point will be around pH = ___. It is important to use the appropriate __________ for each titration - one that has an ____ point (point where it changes colour) near the __________ point.

Ex 1) In a titration, 50.0 mL of 2.00 M HCl is required to reach the endpoint with 25.0 mL of NaOH. Find the concentration of sodium hydroxide.

Ex 2) What is the concentration of HBr if 38.7 mL of 0.0500 M Ca(OH)$_2$ will reach the equivalence point with 20.0 mL?

The Acid Ionization Constant $K_a$ and $pK_a$

$K_a$ is the equilibrium constant for an acid as it ionizes in water. Take the above example:

$$\text{CH}_3\text{COOH}_{(aq)} + \text{H}_2\text{O}_{(l)} \rightleftharpoons \text{H}_3\text{O}^+_{(aq)} + \text{CH}_3\text{COO}^-_{(aq)}$$

$$K_a = \text{Note that H}_2\text{O is left out because it is constant as are all pure solids and liquids: } K_a = K[\text{H}_2\text{O}]$$

The larger the $K_a$, the __________ the acid. An easier way to compare this strength is to use $pK_a$. This works the same way as pH. $pK_a$ has no units and the lower it is, the stronger the acid.

$$pK_a =$$

$K_b$ and $pK_b$

$K_b$ is the equilibrium constant for a base as it reacts with water. In the reaction:

$$\text{NH}_3_{(aq)} + \text{H}_2\text{O}_{(l)} \rightleftharpoons \text{OH}^-_{(aq)} + \text{NH}_4^+_{(aq)} \quad \text{the } K_b =$$

The larger the $K_b$, the ________ the base. As with $pK_a$, $pK_b = $ __________ and the lower the $pK_b$, the stronger the base. Note that $K_a \cdot K_b = $ ______.