

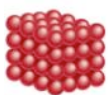



Physical Properties


<p>Gases</p> 	<ul style="list-style-type: none"> • Assumes both the volume and shape of its container • Is compressible • Flows readily • Diffusion within gases occurs rapidly
<p>Liquids</p> 	<ul style="list-style-type: none"> • Assumes the shape of the portion the container it occupies • Does not expand to fill container • Is virtually incompressible • Flows readily • Diffusion within liquids occurs slowly
<p>Solids</p> 	<ul style="list-style-type: none"> • Retains its own shape and volume • Is virtually incompressible • Does not flow • Diffusion within solids occurs EXTREMELY slowly

These properties can be understood in terms of the energy of motion (Kinetic Energy) of the particles of each state compared to the intermolecular forces between those particles. Average Kinetic Energy, which is related to the average speed of particles within any state, is proportional to the absolute temperature of those particles.

Gases-

The average energy of attraction in gas molecules is much smaller than their average kinetic energy. This lack of strong attractive forces allows the gas to expand and fill its container.

Liquids-

In liquids the intermolecular forces are strong enough to hold molecules close together, so liquids are much denser and far less compressible than gases. This gives liquids definite volume. The attractive forces in liquids are not strong enough, however, to keep the molecules from moving past one another. Thus, any liquid can be poured and it assumes the shape of the portion the container it occupies.

Solids-

The intermolecular attractive forces in solids are strong enough not only to hold molecules close together, but also to virtually lock them into place. This means that solids have both definite shape and volume.

Adding heat to solids and liquids gives them the energy to overcome their intermolecular forces and change to a higher energy phase, while removing heat from solids and liquids means that molecules no longer have the energy to overcome those same intermolecular forces and must change to a phase of lower energy.



INTERMOLECULAR FORCES



Ion-Dipole-

Ion-Dipole forces exist between ions and the partial charges on the ends of polar molecules. Polar molecules are dipoles - they have a positive end and negative end due to the differing electronegativities of their component atoms.

Dipole-Dipole-

Neutral polar molecules can attract each other when the positive end of one molecule passes near the negative end of another. This polarity results in increased strength of intermolecular attractions in molecules of relatively equal mass and size.

London Dispersion Forces-

In non-polar molecules, the average distribution of electrons is symmetrical. The molecules are non-polar and possess no permanent dipole moment. The instantaneous distribution of electrons, however, can be different from the average distribution. Because electrons repel one another, the movement of electrons in one atom can influence the motions of the electrons of its neighbors. Thus, the temporary dipole on one atom can induce a similar temporary dipole on an adjacent atom, causing the atoms to be attracted to each other. At any given instant, this process may create an instantaneous dipole moment, giving a momentary polar charge to each end of the molecule. This attractive interaction is called the London Dispersion force. This force is significant only when molecules are close together.

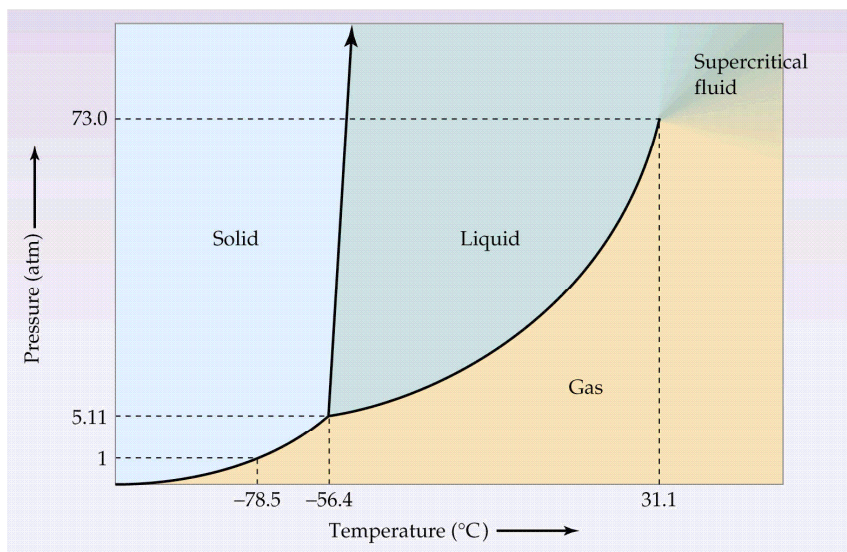
The strength of dispersion forces depends on the ease with which the charge distribution in a molecule can be distorted to induce a momentary dipole. This ease is known as **polarizability**. The shape of molecules also influences the magnitude of dispersion forces. The greater the exposed surface area of a molecule, the greater the chance of a dipole moment attracting another molecule becomes, and the greater the potential attraction is. Substances made up of more massive molecules have stronger dispersion forces.

Hydrogen Bonding-

Hydrogen bonds are a special type of dipole-dipole attraction, and are typically stronger than dispersion forces. They result from the attraction between a hydrogen atom in a polar bond and an unshared electron pair on a nearby, highly electronegative atom such as fluorine, oxygen or nitrogen.

Phase Diagram-

Phase changes depend on the ability of molecules to overcome intermolecular forces. The greater a molecule's kinetic energy, the greater this ability is. A molecule's speed (the result of its KE) is directly related to its temperature. The phase diagram to the right shows how temperature and pressure affect phases.



Quantum Wave Length of Light Ch. 6.1-6.4

6.1

-**Electron Structure**: the arrangement of electrons

-**Electromagnetic Radiation/Radiant energy**: the waves on the electromagnetic spectrum, ranging from gamma rays to radio waves, including the energy we know as visible light

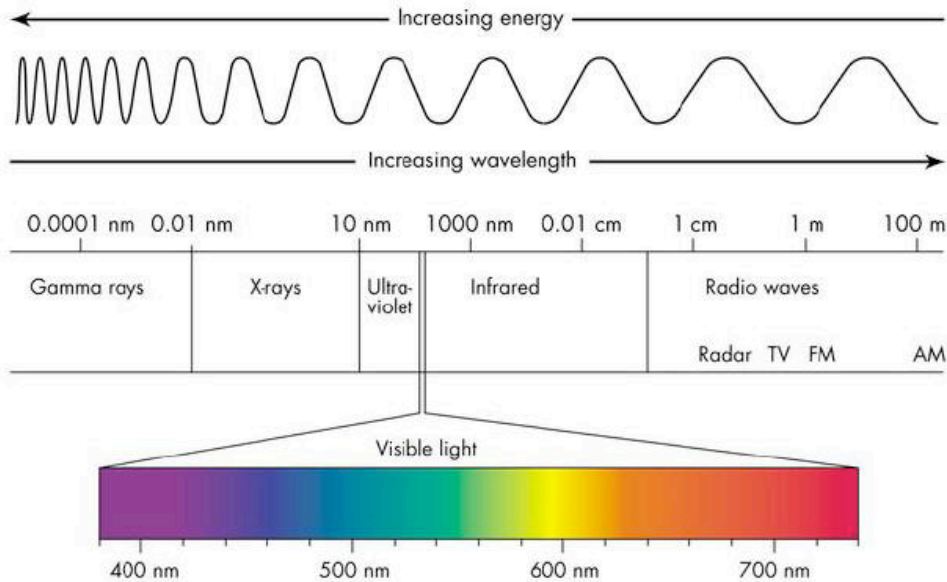
-**Wavelength** is a property of electromagnetic radiation, represented by the symbol λ (lambda), and possessing amplitude A

-**Frequency**, another property, is represented by ν (nu), and is the number of cycles of radiation that pass a point in one second. Frequency is expressed in Hertz ($1 \text{ Hz} = 1 \text{ s}^{-1}$)

-**Speed of light**: $c = \lambda \nu$ $c = 3.00 \times 10^8 \text{ m/s}$

-**Electromagnetic spectrum**: a display of various types of electromagnetic radiation in order of increasing wavelength. See figure 6.4

-**Visible radiation** (to the human eye) is between 400nm and 750nm in wavelength



6.2

-**Blackbody Radiation**: emission of light from hot objects

-**Photoelectric effect**: emission of electrons from metal surfaces exposed to light of certain energies - it provides evidence for particle nature of light and evidence for quantization

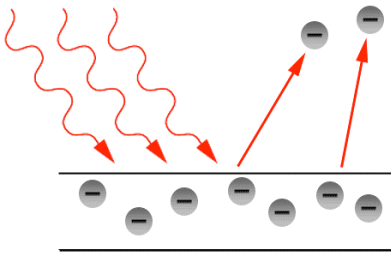
-**Emission spectra**: emissions of light from electronically excited gas atoms

-**Quantum**: smallest amount of energy that can be emitted or absorbed as electromagnetic radiation.

-Relationship between **energy** and **frequency**: $E = h\nu$, where h is **Planck's Constant**, ($6.626 \times 10^{-34} \text{ J-s}$)

-According to Planck's theory, matter is allowed to emit and absorb energy only in whole-number multiples of $h\nu$

-Light is expressed in energy packets called **photons** whose energy is $E = h\nu$



The Photoelectric Effect on metal

6.3

-**Monochromatic**: radiation composed of only one wavelength

-Radiation that spans a whole array of different wavelengths is called **continuous**

-**Spectrum**: When radiation from light is separated into its different wavelength components

-**Continuous Spectrum**: white light separated into colors

-**Line Spectrum**: A spectrum only containing specific wavelengths, not necessarily in consecutive sequence

-Bohr studied the line spectra of certain elements and assumed that electrons were confined to specific energy states called **orbits**, represented by the specific energies of the different frequencies of light

-**Bohr's Model** has three postulates-

1. Only orbits of specific radii, corresponding to certain definite energies, are permitted for electrons in an atom.
2. An electron in a permitted orbit has a specific energy and is in an "allowed" energy state.
3. Energy is only emitted or absorbed by an electron as it moves from one allowed energy state to another. The energy is gained or lost as a photon.

-Amount of energy absorbed or emitted by moving between states:

$$\Delta E = E_f - E_i = h\nu$$

-Limitations of Bohr Model:

1. Can't explain spectra of atoms other than hydrogen.
2. Electrons do not move about the nucleus in circular orbits.

-Important ideas from Bohr Model:

1. Energy of an electron is quantized: electrons exist only in certain energy levels described by quantum numbers.
2. Energy gain or loss is involved in moving an electron from one energy level to another.

6.4

$$-\lambda = h/mv$$

-**Momentum**, mv , is a particle property.

- λ is a wave property.⁸

-**Matter waves**: describe the wave characteristics of material particles.

-**Heisenberg's Uncertainty Principle**: it is not possible to determine the exact position, direction of motion, and speed of subatomic particles simultaneously.

Represented by the equation $\Delta x \cdot \Delta(mv) \geq h/4\pi$

PRACTICE PROBLEMS

1. The orange light given off by a neon light used in 50's style diners has a wavelength of 650nm. What is the frequency of this radiation?

Given: $C = 3.00 \times 10^8$ m/s

Use $C = \nu\lambda$. Solve for ν

$$\lambda = 650\text{nm}$$

$$\nu = ?$$

$$\text{Solve: } \nu = C/\lambda = ((3.00 \times 10^8 \text{m/s}) / (650\text{nm})) \cdot ((1) / (10^{-9}))$$

(Convert nanometers to meters per second for ratio, unit change)

$$\text{Answer} = 4.62 \times 10^{14} \text{s}^{-1}$$

2. Calculate the energy of one photon of blue light whose frequency is $6.74 \times 10^{14} \text{s}^{-1}$.

Use equation $E = h\nu$ to calculate energy.

Use Planck's constant $6.626 \times 10^{-34} \text{J}\cdot\text{s}$ for h and the given frequency for ν .

$$E = (6.626 \times 10^{-34} \text{J}\cdot\text{s})(6.74 \times 10^{14} \text{s}^{-1})$$

$$= 44.65924 \times 10^{-20}$$

$$= 4.465 \times 10^{-19}$$

3. What is the wavelength of an electron moving with a speed of $5.22 \times 10^6 \text{m/s}$ (the mass of the electron is $9.11 \times 10^{-28} \text{g}$)

$$\lambda = h/mv = 6.63 \times 10^{-34} \text{J}\cdot\text{s} / ((9.11 \times 10^{-28})(5.22 \times 10^6 \text{m/s})) \cdot (1 \text{kg m}^2/\text{s}^2) / (1 \text{J}) \cdot (1000 \text{g})$$

$$= 1.39 \times 10^{-10} \text{m} = .139 \text{nm}$$



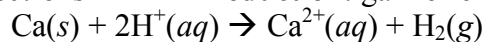
LEO the lion
says GER



Oxidation-reduction reactions: "redox" reactions involving the transfer of electrons between reactants.

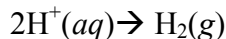
Oxidation: loss of electrons

Reduction: gain of electrons

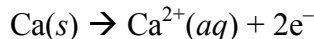


The Ca has lost electrons to gain a 2+ charge, while the 2H^+ has combined to make a diatomic molecule, gaining electrons. The Ca was oxidized and the H^+ was reduced.

The half reaction of the reaction above for reduction:



The half reaction of the reaction above for oxidation:



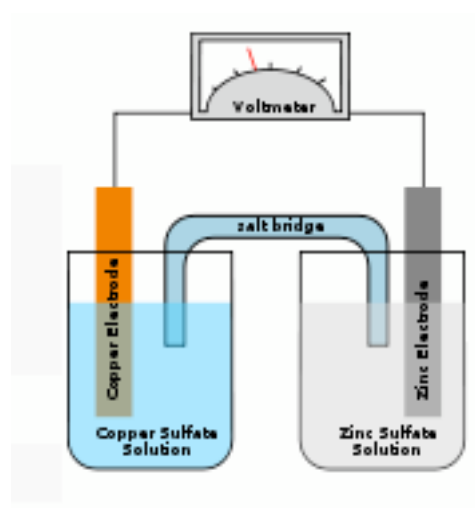
Oxidation numbers:

1. An atom in its elemental form has an oxidation number of zero (Example: H_2)
2. Monatomic ions has an oxidation number that equals their charge (Example: $\text{K}^+ = 1+$ oxidation number), their oxidation number will be shown by their group number
3. Nonmetals usually have negative oxidation numbers (Example: N has a 3- oxidation number)
4. The sum of the oxidation numbers in a compound equals the charge of the compound (Example: NaCl is 0)

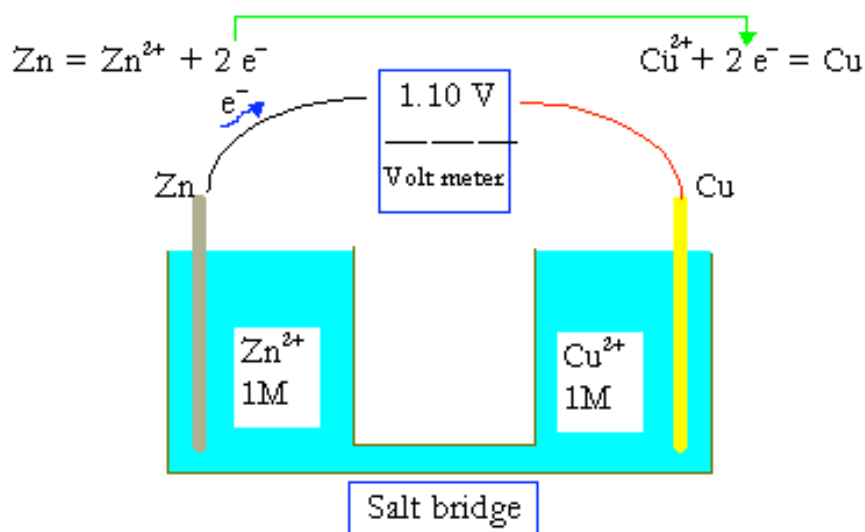
Remember-

- Oxygen is usually 2- except when forming O_2^{2-} where it is 1-
- Hydrogen is always 1+ when bonded to nonmetals, and 1- when bonded metals
- Fluorine is 1- in all compounds
- Halogens are 1- except when combined with oxygen, then they are 1+
- If the oxidation number increase then it is oxidation, if the oxidation number decrease then it is reduction
- Oxygen is always the oxidizing agent, it is always reduced (unless it bonds with fluorine)
- Corrosion: oxidation of metals to their oxides
- Whatever is reduced is the oxidizing agent, because it makes the other thing get oxidized

Activity Series: This chart can be used to determine the voltage of a voltaic cell. The anode of the cell will be oxidized and the cathode will be reduced, causing the cathode's electrode to gain mass and the anode's to shrink.



- The anode is on the left, and the cathode is on the right.
- The anode is the anion and the cathode is the cation, in the half-reactions.
- There is always one mole of each half cell.
- Electrons are being transferred through the voltmeter and are being stabilized by the salt bridge.
- Use the half-reactions from the activity series, the anode minus the cathode, in order to determine the voltage of the battery, or standard cell potential
- The concentration is always a 1:1 ratio of the half cells



Anode: Cu^{2+} Cathode: Zn^2 Standard cell potential: $Zn = Zn^{2+} + 2e^-$ ($V = -0.76$) (oxidized) and $Cu^{2+} + 2e^- = Cu$ ($V = 0.34$) (reduced) $0.34 - (-0.76) = 1.10 V$

Standard Reduction Potentials (25°C)

Half-Cell Reactions	E°
$\text{F}_{2(\text{g})} + 2\text{e}^{-1} \longrightarrow 2\text{F}^{-1}_{(\text{aq})}$	+2.87
$\text{O}_{3(\text{g})} + 2\text{H}^{+1}_{(\text{aq})} + 2\text{e}^{-1} \longrightarrow \text{O}_{2(\text{g})} + \text{H}_2\text{O}_{(\text{l})}$	+2.08
$\text{S}_2\text{O}_8^{2-}_{(\text{aq})} + 2\text{e}^{-1} \longrightarrow 2\text{SO}_4^{2-}_{(\text{aq})}$	+2.05
$\text{Co}^{3+}_{(\text{aq})} + \text{e}^{-1} \longrightarrow \text{Co}^{2+}_{(\text{aq})}$	+1.82
$\text{H}_2\text{O}_{2(\text{aq})} + 2\text{H}^{+1}_{(\text{aq})} + 2\text{e}^{-1} \longrightarrow 2\text{H}_2\text{O}_{(\text{l})}$	+1.77
$\text{MnO}_4^{-1}_{(\text{aq})} + 4\text{H}^{+1}_{(\text{aq})} + 3\text{e}^{-1} \longrightarrow \text{MnO}_{2(\text{s})} + 2\text{H}_2\text{O}_{(\text{l})}$	+1.695
$\text{PbO}_{2(\text{s})} + \text{SO}_4^{2-}_{(\text{aq})} + 4\text{H}^{+1}_{(\text{aq})} + 2\text{e}^{-1} \longrightarrow \text{PbSO}_{4(\text{s})} + 2\text{H}_2\text{O}_{(\text{l})}$	+1.69
$2\text{HOCl}_{(\text{aq})} + 2\text{H}^{+1}_{(\text{aq})} + 2\text{e}^{-1} \longrightarrow \text{Cl}_{2(\text{g})} + 2\text{H}_2\text{O}_{(\text{l})}$	+1.63
$\text{Mn}^{3+}_{(\text{aq})} + \text{e}^{-1} \longrightarrow \text{Mn}^{2+}_{(\text{aq})}$	+1.51
$\text{MnO}_4^{-1}_{(\text{aq})} + 8\text{H}^{+1}_{(\text{aq})} + 5\text{e}^{-1} \longrightarrow \text{Mn}^{2+}_{(\text{aq})} + 4\text{H}_2\text{O}_{(\text{l})}$	+1.49
$\text{PbO}_{2(\text{s})} + 4\text{H}^{+1}_{(\text{aq})} + 2\text{e}^{-1} \longrightarrow \text{Pb}^{2+}_{(\text{aq})} + 2\text{H}_2\text{O}_{(\text{l})}$	+1.46
$\text{BrO}_3^{-1}_{(\text{aq})} + 6\text{H}^{+1}_{(\text{aq})} + 6\text{e}^{-1} \longrightarrow \text{Br}^{-1}_{(\text{aq})} + 3\text{H}_2\text{O}_{(\text{l})}$	+1.44
$\text{Au}^{3+}_{(\text{aq})} + 3\text{e}^{-1} \longrightarrow \text{Au}_{(\text{s})}$	+1.42
$\text{Cl}_{2(\text{g})} + 2\text{e}^{-1} \longrightarrow 2\text{Cl}^{-1}_{(\text{aq})}$	+1.36
$\text{Cr}_2\text{O}_7^{2-}_{(\text{aq})} + 14\text{H}^{+1}_{(\text{aq})} + 6\text{e}^{-1} \longrightarrow 2\text{Cr}^{3+}_{(\text{aq})} + 7\text{H}_2\text{O}_{(\text{l})}$	+1.33
$\text{O}_{3(\text{g})} + \text{H}_2\text{O}_{(\text{l})} + 2\text{e}^{-1} \longrightarrow \text{O}_{2(\text{g})} + 2\text{OH}^{-1}_{(\text{aq})}$	+1.24
$\text{MnO}_{2(\text{s})} + 4\text{H}^{+1}_{(\text{aq})} + 2\text{e}^{-1} \longrightarrow \text{Mn}^{2+}_{(\text{aq})} + 2\text{H}_2\text{O}_{(\text{l})}$	+1.23
$\text{O}_{2(\text{g})} + 4\text{H}^{+1}_{(\text{aq})} + 4\text{e}^{-1} \longrightarrow 2\text{H}_2\text{O}_{(\text{l})}$	+1.23
$\text{Pt}^{2+}_{(\text{aq})} + 2\text{e}^{-1} \longrightarrow \text{Pt}_{(\text{s})}$	+1.20
$\text{Br}_{2(\text{aq})} + 2\text{e}^{-1} \longrightarrow 2\text{Br}^{-1}_{(\text{aq})}$	+1.07
$\text{NO}_3^{-1}_{(\text{aq})} + 4\text{H}^{+1}_{(\text{aq})} + 3\text{e}^{-1} \longrightarrow \text{NO}_{(\text{g})} + 2\text{H}_2\text{O}_{(\text{l})}$	+0.96
$\text{NO}_3^{-1}_{(\text{aq})} + 3\text{H}^{+1}_{(\text{aq})} + 2\text{e}^{-1} \longrightarrow \text{HNO}_{2(\text{g})} + \text{H}_2\text{O}_{(\text{l})}$	+0.94
$2\text{Hg}^{2+}_{(\text{aq})} + 2\text{e}^{-1} \longrightarrow \text{Hg}_2^{2+}_{(\text{aq})}$	+0.91
$\text{HO}_2^{-1}_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})} + 2\text{e}^{-1} \longrightarrow 3\text{OH}^{-1}_{(\text{aq})}$	+0.87
$2\text{NO}_3^{-1}_{(\text{aq})} + 4\text{H}^{+1}_{(\text{aq})} + 2\text{e}^{-1} \longrightarrow 2\text{NO}_{2(\text{g})} + 2\text{H}_2\text{O}_{(\text{l})}$	+0.80
$\text{Ag}^{+1}_{(\text{aq})} + \text{e}^{-1} \longrightarrow \text{Ag}_{(\text{s})}$	+0.80
$\text{Fe}^{3+}_{(\text{aq})} + \text{e}^{-1} \longrightarrow \text{Fe}^{2+}_{(\text{aq})}$	+0.77
$\text{O}_{2(\text{g})} + 2\text{H}^{+1}_{(\text{aq})} + 2\text{e}^{-1} \longrightarrow \text{H}_2\text{O}_{2(\text{aq})}$	+0.69
$\text{I}_{2(\text{s})} + 2\text{e}^{-1} \longrightarrow 2\text{I}^{-1}_{(\text{aq})}$	+0.54
$\text{NiO}_{2(\text{s})} + 2\text{H}_2\text{O}_{(\text{l})} + 2\text{e}^{-1} \longrightarrow \text{Ni}(\text{OH})_2 + 2\text{OH}^{-1}_{(\text{aq})}$	+0.49
$\text{SO}_{2(\text{aq})} + 4\text{H}^{+1}_{(\text{aq})} + 4\text{e}^{-1} \longrightarrow \text{S}_{(\text{s})} + 2\text{H}_2\text{O}_{(\text{l})}$	+0.45
$\text{O}_{2(\text{g})} + 2\text{H}_2\text{O}_{(\text{l})} + 4\text{e}^{-1} \longrightarrow 4\text{OH}^{-1}_{(\text{aq})}$	+0.401

$\text{Cu}^{2+}_{(\text{aq})} + 2\text{e}^{-1} \text{-----} \rightarrow \text{Cu}_{(\text{s})}$	+0.34
$\text{Hg}_2\text{Cl}_{2(\text{s})} + 2\text{e}^{-1} \text{-----} \rightarrow 2 \text{Hg}_{(\text{l})} + 2 \text{Cl}^{-1}_{(\text{aq})}$	+0.27
$\text{PbO}_{2(\text{s})} + \text{H}_2\text{O}_{(\text{l})} + 2\text{e}^{-1} \text{-----} \rightarrow \text{PbO}_{(\text{s})} + 2 \text{OH}^{-1}_{(\text{aq})}$	+0.25
$\text{AgCl}_{(\text{s})} + \text{e}^{-1} \text{-----} \rightarrow \text{Ag}_{(\text{s})} + \text{Cl}^{-1}_{(\text{aq})}$	+0.2223
$\text{SO}_4^{2-}_{(\text{aq})} + 4\text{H}^{+1}_{(\text{aq})} + 2\text{e}^{-1} \text{-----} \rightarrow \text{H}_2\text{SO}_{3(\text{aq})} + \text{H}_2\text{O}_{(\text{l})}$	+0.172
$\text{S}_4\text{O}_6^{2-}_{(\text{aq})} + 2\text{e}^{-1} \text{-----} \rightarrow 2 \text{S}_2\text{O}_3^{2-}_{(\text{aq})}$	+0.169
$\text{Cu}^{2+}_{(\text{aq})} + \text{e}^{-1} \text{-----} \rightarrow \text{Cu}^{+1}_{(\text{aq})}$	+0.16
$\text{Sn}^{4+}_{(\text{aq})} + 2\text{e}^{-1} \text{-----} \rightarrow \text{Sn}^{2+}_{(\text{aq})}$	+0.15
$\text{S}_{(\text{s})} + 2\text{H}^{+1}_{(\text{aq})} + 2\text{e}^{-1} \text{-----} \rightarrow \text{H}_2\text{S}_{(\text{g})}$	+0.14
$\text{AgBr}_{(\text{s})} + \text{e}^{-1} \text{-----} \rightarrow \text{Ag}_{(\text{s})} + \text{Br}^{-1}_{(\text{aq})}$	+0.07
$2 \text{H}^{+1}_{(\text{aq})} + 2\text{e}^{-1} \text{-----} \rightarrow \text{H}_{2(\text{g})}$	0.00
$\text{Pb}^{2+}_{(\text{aq})} + 2\text{e}^{-1} \text{-----} \rightarrow \text{Pb}_{(\text{s})}$	-0.13
$\text{Sn}^{2+}_{(\text{aq})} + 2\text{e}^{-1} \text{-----} \rightarrow \text{Sn}_{(\text{s})}$	-0.14
$\text{AgI}_{(\text{s})} + \text{e}^{-1} \text{-----} \rightarrow \text{Ag}_{(\text{s})} + \text{I}^{-1}_{(\text{aq})}$	-0.15
$\text{Ni}^{2+}_{(\text{aq})} + 2\text{e}^{-1} \text{-----} \rightarrow \text{Ni}_{(\text{s})}$	-0.25
$\text{Co}^{2+}_{(\text{aq})} + 2\text{e}^{-1} \text{-----} \rightarrow \text{Co}_{(\text{s})}$	-0.28
$\text{In}^{3+}_{(\text{aq})} + 3\text{e}^{-1} \text{-----} \rightarrow \text{In}_{(\text{s})}$	-0.34
$\text{Tl}^{+1}_{(\text{aq})} + \text{e}^{-1} \text{-----} \rightarrow \text{Tl}_{(\text{s})}$	-0.34
$\text{PbSO}_{4(\text{s})} + 2\text{e}^{-1} \text{-----} \rightarrow \text{Pb}_{(\text{s})} + \text{SO}_4^{2-}_{(\text{aq})}$	-0.36
$\text{Cd}^{2+}_{(\text{aq})} + 2\text{e}^{-1} \text{-----} \rightarrow \text{Cd}_{(\text{s})}$	-0.40
$\text{Fe}^{2+}_{(\text{aq})} + 2\text{e}^{-1} \text{-----} \rightarrow \text{Fe}_{(\text{s})}$	-0.44
$\text{Ga}^{3+}_{(\text{aq})} + 3\text{e}^{-1} \text{-----} \rightarrow \text{Ga}_{(\text{s})}$	-0.56
$\text{PbO}_{(\text{s})} + \text{H}_2\text{O}_{(\text{l})} + 2\text{e}^{-1} \text{-----} \rightarrow \text{Pb}_{(\text{s})} + 2 \text{OH}^{-1}_{(\text{aq})}$	-0.58
$\text{Cr}^{3-}_{(\text{aq})} + 3\text{e}^{-1} \text{-----} \rightarrow \text{Cr}_{(\text{s})}$	-0.74
$\text{Zn}^{2+}_{(\text{aq})} + 2\text{e}^{-1} \text{-----} \rightarrow \text{Zn}_{(\text{s})}$	-0.76
$\text{Cd}(\text{OH})_{2(\text{s})} + 2\text{e}^{-1} \text{-----} \rightarrow \text{Cd}_{(\text{s})} + 2 \text{OH}^{-1}_{(\text{aq})}$	-0.81
$2 \text{H}_2\text{O}_{(\text{l})} + 2\text{e}^{-1} \text{-----} \rightarrow \text{H}_{2(\text{g})} + 2 \text{OH}^{-1}_{(\text{aq})}$	-0.83
$\text{Fe}(\text{OH})_{2(\text{s})} + 2\text{e}^{-1} \text{-----} \rightarrow \text{Fe}_{(\text{s})} + 2 \text{OH}^{-1}_{(\text{aq})}$	-0.88
$\text{Cr}^{2+}_{(\text{aq})} + 2\text{e}^{-1} \text{-----} \rightarrow \text{Cr}_{(\text{s})}$	-0.91
$\text{N}_{2(\text{g})} + 4 \text{H}_2\text{O}_{(\text{l})} + 4\text{e}^{-1} \text{-----} \rightarrow \text{N}_2\text{O}_{4(\text{aq})} + 4 \text{OH}^{-1}_{(\text{aq})}$	-1.16
$\text{V}^{2+}_{(\text{aq})} + 2\text{e}^{-1} \text{-----} \rightarrow \text{V}_{(\text{s})}$	-1.18
$\text{ZnO}_2^{-1}_{(\text{aq})} + 2 \text{H}_2\text{O}_{(\text{l})} + 2\text{e}^{-1} \text{-----} \rightarrow \text{Zn}_{(\text{s})} + 4\text{OH}^{-1}_{(\text{aq})}$	-1.216
$\text{Ti}^{2+}_{(\text{aq})} + 2\text{e}^{-1} \text{-----} \rightarrow \text{Ti}_{(\text{s})}$	-1.63
$\text{Al}^{3+}_{(\text{aq})} + 3\text{e}^{-1} \text{-----} \rightarrow \text{Al}_{(\text{s})}$	-1.66
$\text{U}^{3+}_{(\text{aq})} + 3\text{e}^{-1} \text{-----} \rightarrow \text{U}_{(\text{s})}$	-1.79

$\text{Sc}^{3+}_{(\text{aq})} + 3\text{e}^{-1} \text{-----} \rightarrow \text{Sc}_{(\text{s})}$	-2.02
$\text{La}^{3+}_{(\text{aq})} + 3\text{e}^{-1} \text{-----} \rightarrow \text{La}_{(\text{s})}$	-2.36
$\text{Y}^{3+}_{(\text{aq})} + 3\text{e}^{-1} \text{-----} \rightarrow \text{Y}_{(\text{s})}$	-2.37
$\text{Mg}^{2+}_{(\text{aq})} + 2\text{e}^{-1} \text{-----} \rightarrow \text{Mg}_{(\text{s})}$	-2.37
$\text{Na}^{+1}_{(\text{aq})} + \text{e}^{-1} \text{-----} \rightarrow \text{Na}_{(\text{s})}$	-2.71
$\text{Ca}^{2+}_{(\text{aq})} + 2\text{e}^{-1} \text{-----} \rightarrow \text{Ca}_{(\text{s})}$	-2.76
$\text{Sr}^{2+}_{(\text{aq})} + 2\text{e}^{-1} \text{-----} \rightarrow \text{Sr}_{(\text{s})}$	-2.89
$\text{Ba}^{2+}_{(\text{aq})} + 2\text{e}^{-1} \text{-----} \rightarrow \text{Ba}_{(\text{s})}$	-2.90
$\text{Cs}^{+1}_{(\text{aq})} + \text{e}^{-1} \text{-----} \rightarrow \text{Cs}_{(\text{s})}$	-2.92
$\text{K}^{+1}_{(\text{aq})} + \text{e}^{-1} \text{-----} \rightarrow \text{K}_{(\text{s})}$	-2.92
$\text{Rb}^{+1}_{(\text{aq})} + \text{e}^{-1} \text{-----} \rightarrow \text{Rb}_{(\text{s})}$	-2.93
$\text{Li}^{+1}_{(\text{aq})} + \text{e}^{-1} \text{-----} \rightarrow \text{Li}_{(\text{s})}$	-3.05

Types of reactions:

Combustion: $\text{CH}_4 + 4\text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$

A fuel plus oxygen yields water and (usually) carbon dioxide.

Synthesis: $2\text{H}_2 + \text{C} \rightarrow \text{CH}_4$

Multiple reactants form a single product.

Decomposition: $2\text{H}_2\text{O} \rightarrow 2\text{H}_2 + \text{O}_2$

A reactant breaks down into multiple products.

Double displacement: $\text{KCl} + \text{AgNO}_3 \rightarrow \text{AgCl}_{(\text{s})} + \text{KNO}_3$

Ions exchange "partners."

Single displacement: $\text{Cu}(\text{NO}_3)_2 + \text{Zn}_{(\text{s})} \rightarrow \text{Cu}_{(\text{s})} + \text{Zn}(\text{NO}_3)_2$

The compound reactant switches with the lone reactant.

Redox: $\text{H}_2 + \text{F}_2 \rightarrow 2 \text{HF} = (\text{H}_2 \rightarrow 2 \text{H}^+ + 2 \text{e}^-) + (\text{F}_2 + 2 \text{e}^- \rightarrow 2 \text{F}^-)$

Transfer of electrons causes oxidation and reduction. (LEO the lion says GER) Here, hydrogen is reduced and fluorine is oxidized.

Precipitation: $\text{AgNO}_3 (\text{aq}) + \text{KCl} (\text{aq}) \rightarrow \text{AgCl} (\text{s}) + \text{KNO}_3 (\text{aq})$

Two aqueous solutions react to form a solid that usually falls out of solution.

Anhydride: $(\text{CH}_3\text{CO})_2\text{O} + \text{H}_2\text{O} \rightarrow 2\text{CH}_3\text{COOH}$

Reactant (an anhydride) plus water forms an acid.

Complex ion formation: $\text{Ag}^+_{(\text{aq})} + 2\text{NH}_3 \rightarrow \text{Ag}(\text{NH}_3)_2^+$

A complex ion contains a metal cation bonded to one or more small molecules or ions.

Net Ionic Equations:

Soluble salts and strong acids and bases will completely ionize in solution, so in an equation they are written as ions.

Weak acids and bases, molecular compounds, insoluble salts, solids, and gases will not ionize significantly.

To write a net ionic equation, eliminate all spectator ions (those who remain unchanged by the reaction) from the equation.

Strong Acids and Bases:

<u>Acids</u> pH less than 7 turns blue litmus paper red donates hydrides (H^+) tastes sour corrosive to metals	<u>Bases</u> pH greater than 7 turns red litmus paper blue accepts hydrides / usually has hydroxides (OH^-) tastes bitter slippery
---	---

Strong acids and bases are electrolytes, and have weak conjugates. Weak acids and bases have strong conjugates.

<u>Strong Acids</u> HCl HBr HI H ₂ SO ₄ HClO ₃ HClO ₄ HNO ₃	<u>Strong Bases</u> All group 1A metal hydroxides Ca(OH) ₂ Ba(OH) ₂ Sr(OH) ₂
<u>Weak Acids</u> CH ₃ COOH HCOOH HF H ₂ S HCN NH ₄ ⁺	<u>Weak Bases</u> NH ₃ C ₅ H ₅ N NH ₄ OH

Water can be considered both a weak acid and a weak base because it can give or accept a hydride.

Solubility Rules:

1. Salts containing Group I elements are soluble (Li^+ , Na^+ , K^+ , Cs^+ , Rb^+). Exceptions to this rule are rare. Salts containing the ammonium ion (NH_4^+) are also soluble.
2. Salts containing the nitrate ion (NO_3^-) are generally soluble.
3. Salts containing Cl^- , Br^- , I^- are generally soluble. Important exceptions to this rule are halide salts of Ag^+ , Pb^{2+} , and $(Hg_2)^{2+}$. Thus, $AgCl$, $PbBr_2$, and Hg_2Cl_2 are all insoluble.
4. Most silver salts are insoluble. $AgNO_3$ and $Ag(C_2H_3O_2)$ are common soluble salts of silver; virtually anything else is insoluble.
5. Most sulfate salts are soluble. Important exceptions to this rule include $BaSO_4$, $PbSO_4$, Ag_2SO_4 and $SrSO_4$.
6. Most hydroxide salts are only slightly soluble. Hydroxide salts of Group I elements are soluble. Hydroxide salts of Group II elements (Ca, Sr, and Ba) are slightly soluble. Hydroxide salts of transition metals and Al^{3+} are insoluble. Thus, $Fe(OH)_3$, $Al(OH)_3$, $Co(OH)_2$ are not soluble.
7. Most sulfides of transition metals are highly insoluble. Thus, CdS , FeS , ZnS , Ag_2S are all insoluble. Arsenic, antimony, bismuth, and lead sulfides are also insoluble.
8. Carbonates are frequently insoluble. Group II carbonates (Ca, Sr, and Ba) are insoluble. Some other insoluble carbonates include $FeCO_3$ and $PbCO_3$.

9. Chromates are frequently insoluble. Examples: PbCrO_4 , BaCrO_4
 10. Phosphates are frequently insoluble. Examples: $\text{Ca}_3(\text{PO}_4)_2$, Ag_3PO_4
 11. Fluorides are frequently insoluble. Examples: BaF_2 , MgF_2 , PbF_2 .

Thermodynamics:

The SI unit of energy is the **joule (J)**. 1 calorie = 4.184 J

Nomenclature:

-Organic Molecules:

- Hydrocarbons are the simplest form of organic molecules, made of only hydrogen and carbon.
- If carbon has all single bonds, its name always ends in “ane”.
- The beginning of a molecule depends on the number of molecules in the chain.

Suffixes:

# of Carbons:	Prefixes:	w/ single bond	w/ double bond	w/ triple bond
1	Meth-	-ane	-----	-----
2	Eth-	-ane	-ene	-yne
3	Prop-	-ane	-ene	-yne
4	But-	-ane	-ene	-yne
5	Pent-	-ane	-ene	-yne
6	Hex-	-ane	-ene	-yne
7	Hept-	-ane	-ene	-yne

-If there is an OH functional group present in the molecule replacing an H it is an alcohol and therefore ends in “ol”. CH_3OH is *methanol*.

-If there is an OOH functional group replacing the H in the molecule it is an acid, and it is called prefix + oic + acid. HCOOH is *methanoic acid* (or formic acid)

-Inorganic Compounds:

-**Ionic:** Name of the cation + name of the anion. The name of the anion is the element + “ide.” NaCl is Sodium (cation) + Chloride (anion). When it is ionic the numerical prefix isn't stated. MgF_2 is Magnesium fluoride, **NOT** Magnesium difluoride (wrong).

-Polyatomic Ions:

TABLE 4.4

Common Polyatomic Ions					
1- charge		2- charge		3- charge	
Formula	Name	Formula	Name	Formula	Name
H_2PO_4^-	Dihydrogen phosphate	HPO_4^{2-}	Hydrogen phosphate	PO_3^{3-}	Phosphite
$\text{C}_2\text{H}_3\text{O}_2^-$	Acetate	$\text{C}_2\text{O}_4^{2-}$	Oxalate	PO_4^{3-}	Phosphate
HSO_3^-	Hydrogen sulfite	SO_3^{2-}	Sulfite		
HSO_4^-	Hydrogen sulfate	SO_4^{2-}	Sulfate		
HCO_3^-	Hydrogen carbonate	CO_3^{2-}	Carbonate		
NO_2^-	Nitrite	CrO_4^{2-}	Chromate		
NO_3^-	Nitrate	$\text{Cr}_2\text{O}_7^{2-}$	Dichromate		
CN^-	Cyanide	SiO_3^{2-}	Silicate		
OH^-	Hydroxide				
MnO_4^-	Permanganate				
ClO^-	Hypochlorite				
ClO_2^-	Chlorite				
ClO_3^-	Chlorate				
ClO_4^-	Perchlorate				

- **Covalent:** For binary (2 elements) compounds, nomenclature is the same as ionic, but the numerical prefix is stated. Mono is used only for the anion. CO is Carbon monoxide. N₂O₅ is Dinitrogen pentoxide.

- For more complex molecules, refer to the organic molecules table.

- Acids:

- Most acids consist of one or more hydrogen atoms bonded to an anion, excluding the insanely complicated and organic ones.

H

+ monatomic ion= hydro+_____ ic acid.

H₂S is hydrosulfuric acid.

+ polyatomic ion ending in -ate= _____ ic acid.

(Note: No hydro)

H₂SO₄ is sulfuric acid.

+ polyatomic ion ending in -ite= _____ ous acid.

(Note: No hydro)

H₂SO₃ is sulfurous acid.

-Miscellaneous:

-Roman Numerals denote the charge of the elements ion. Copper (II) is Cu²⁺ and Copper (I) is Cu⁺. Mercury is an exception- (Hg₂)²⁺ is Mercury (I) because each molecule has a charge of 1⁺.

- **Some Tricky Elements:**

Sodium is Na

Potassium is K

Tin is Sn

Lead is Pb

Gold is Au

Silver is Ag

Iron is Fe

Tungsten is W

Antimony is Sb

Mercury is Hg

Arsenic is As

Cobalt is Co v. Copper is Cu

Magnesium is Mg v. Manganese is Mn

-**Pressure:** The idea of a force that tends to move something: $P = \frac{F}{A}$ where F = force = ma (Mass times acceleration) and A=area.

The SI unit of force is the **newton** (N).

•1 N = 1 kg·m/s²

•The SI unit of pressure is the **pascal** (Pa).

•1 Pa = 1 N/m²

•A related unit is the **bar**, which is equal to 10⁵ Pa.

•1 atm = 760 mm Hg = 760 torr = 1.01325 x 10⁵ Pa = 101.325 kPa

•**Standard atmospheric pressure** is the pressure required to support 760 mm of Hg in a barometer.

Boyle's Law- the volume of a fixed quantity of gas maintained at constant temperature is inversely proportional to the pressure. (when you compress a gas, its volume gets smaller)

PV= constant

Charles's Law- the volume of a fixed amount of gas is maintained at constant pressure is directly proportional to its absolute temperature.

$$V = \text{constant} \times T \text{ or } V/T = \text{constant}$$

Avogadro's Law- the volume of a gas maintained at constant temperature and pressure is directly proportional to the number of moles of the gas.

$$V = \text{constant} \times n$$

Ideal Gas Equation- $PV = nRT$

R= the gas constant (0.08206 L-atm/ mol-K)

n= the # of moles

P= pressure in atm

V=volume in Liters

Standard Temperature and Pressure (STP)= 0° C and 1 atm

Molar volume of an ideal gas at STP is 22.4 L

Relating the Ideal Gas Equation

When any of the variables P, V, n, or T are held constant, they are grouped together and labeled under one name, constant. By doing this, we can form all of the gas laws from $PV = nRT$

Ex. When the quantity of gas and the temperature are held constant, n and T have fixed values.

Therefore nRT is a constant. (Boyle's Law)

$PV = nRT = \text{constant}$ or $PV = \text{constant}$

•Calculating gas density:

$$\therefore d = \frac{PM}{RT}$$

•The molar mass of a gas can be determined as follows: $M = \frac{dRT}{P}$

-Dalton's Law: Total pressure of a mixture of gases equals the sum of the pressures that each would exert if they were present alone. $P_t = P_1 + P_2 + P_3 + \dots$

-**Partial Pressure**: Pressure exerted by a particular component of a mixture of gases.

The total pressure at constant temperature and volume is determined by the number of moles of gas present.

$$P_t = (n_1 + n_2 + n_3 + \dots) \frac{RT}{V} = n_t \frac{RT}{V}$$

-**Mole Fractions**: (n_1/n_t).

Mole fractions are the ratio of the number of moles of one component to the total number of moles in the mixture. Partial pressure equals the mole fraction times the total pressure.

-**Pressure of gases over water**: $P_{\text{total}} = P_{\text{gas}} + P_{\text{water}}$

KINETIC MOLECULAR THEORY

•Root-mean-square (rms) speed, u , is the speed of a gas molecule having average kinetic energy.

•Average kinetic energy, e , is related to rms speed by the equation:

$$e = \frac{1}{2}mu^2 \quad \text{where } m = \text{mass of the molecule.}$$

-With the same temperature, two gases, regardless of mass, will have equal average kinetic energy.

-When u =rms speed and M =molar mass, then $\frac{r_1}{r_2} = \frac{u_1}{u_2} = \sqrt{\frac{M_2}{M_1}}$

-**Effusion**: escape of gas particles, directly proportional to the square root of molar mass.

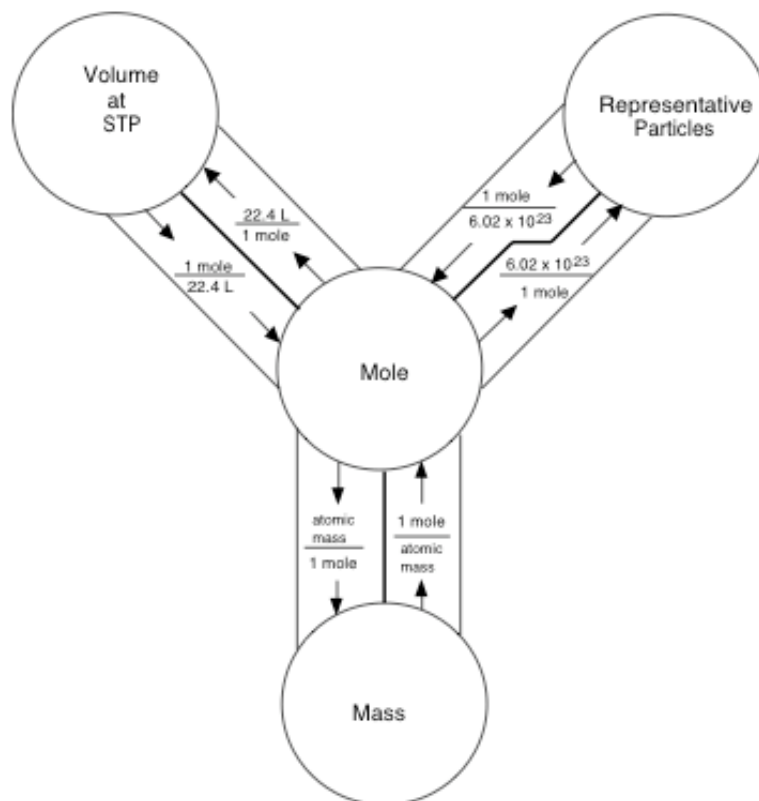
$$u = \sqrt{\frac{3RT}{M}}$$

-**Diffusion**: spread of a substance through space, or a second substance

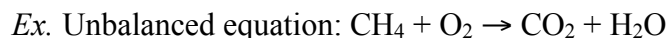
Stoichiometry, And A Million Ways To Make A Mole

-To do stoichiometry, you need moles.

Mole Road Map:



- **Law of Conservation Of Mass**: same number of atoms (and moles) on each side of the equation.

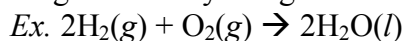


- The **mole ratio** is the number of moles of a product or reactant compared to another product or reactant.

Ex. The mole ratio of Oxygen to Carbon Dioxide is 2 O_2 : 1 CO_2 . (It takes 2 moles of Oxygen reacted with excess Methane to produce 1 mole of Carbon dioxide.)

$$\text{Mole fraction of component, } X = \frac{\text{moles of component}}{\text{total moles of all components}}$$

- The reactants that are completely consumed are called the **limiting reactants or limiting reagents**. (You find the limiting reactant by using the mole ratio.)



- The balanced chemical equation tells us that the stoichiometric ratio of H₂ to O₂ is 2 to 1.
- This means that our 10 H₂ molecules require 5 O₂ molecules (2:1).
- Since we have 7 O₂ molecules, our reaction is *limited* by the amount of H₂ we have (the O₂ is in excess).
- So, all 10 H₂ molecules can (and do) react with 5 of the O₂ molecules producing 10 H₂O molecules.
- At the end of the reaction, 2 O₂ molecules remain unreacted.

Molar Mass

- The **molar mass** is how many grams are in 1 mole of a given substance.
- The **molar mass** has the same value as the atomic mass.

11	Atomic number
Na	Element symbol
Sodium	Element name
22.99	Average atomic mass*

Ex. 1 mole of Sodium is 22.99 grams.

- To find the experimental molar mass of a gas (using data) you find the volume of the gas at STP and convert that volume to moles using the molar road map. Take the measured weight of the gas and divide by the number of moles you just calculated.
- Another way is to use the density of a gas (grams/Liters) and multiply it by 22.4L/mol (molar volume of a gas at STP).
- **Molality** does not change with temperature. Molality, $m = \frac{\text{moles of solute}}{\text{kilograms of solvent}}$
- To find molar mass with molality, you multiply the molality by Kg of solvent, giving you the moles of solute. Then you take the grams of solute and divide it by the moles you just found.
- Multiply **molarity** by volume to get moles of solute, then divide grams of solute by the moles calculated. Molarity, $M = \frac{\text{moles of solute}}{\text{liters of solution}}$

Molar Mass From Boiling Point Elevation

1. Determine the change in boiling point from the observed boiling point of the solution and the boiling point of the pure solvent.

$$\Delta T = T_{\text{solution}} - T_{\text{pure solvent}}$$

2. Determine the molal concentration from the change in boiling point and the boiling point elevation constant.

$$\Delta T = K_b m \quad \text{or} \quad m = \frac{\Delta T}{K_b}$$

3. Determine the moles of unknown (the solute) from the molality of the solution and the mass of solvent (in kilograms) used to make the solution.

$$\text{Molality} = \frac{\text{moles of solute}}{\text{mass of solvent in kilograms}} \quad \text{Moles of solute} = \text{molality} \times \text{mass of solvent}$$

4. Determine the molar mass from the mass of the unknown and the number of moles of unknown.

$$\text{Moles of unknown} = \frac{\text{mass of unknown}}{\text{molar mass}} \quad \text{Molar mass} = \frac{\text{mass of unknown}}{\text{moles of unknown}}$$

Molar Mass From Freezing Point Depression

1. Determine the change in freezing point from the observed freezing point of the solution and the freezing point of the pure solvent. (This will be a negative number.)

$$\Delta T = T_{\text{solution}} - T_{\text{pure solvent}}$$

2. Determine the molal concentration, from the change in freezing point and the freezing point depression constant.

$$\Delta T = -K_f m \quad m = \frac{\Delta T}{-K_f}$$

3. Follow steps 3 and 4 of Molar Mass From Boiling Point Elevation.

Molar Mass From Osmotic Pressure

1. Determine the molar concentration of the unknown in the solution from the observed osmotic pressure.

$$\pi = MRT \quad M = \frac{\pi}{RT}$$

2. Determine the moles of unknown (the solute) from the molarity of the solution and the volume (in liters) of the solution.

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{volume of solution in liters}}$$

Moles of solute = Molarity \times volume of solution in liters

3. Determine the molar mass from the mass of the unknown and the number of moles of unknown.


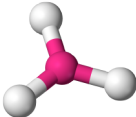
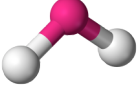
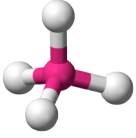
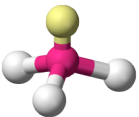
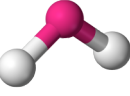
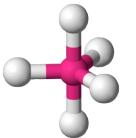
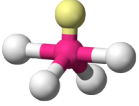
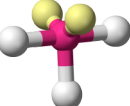
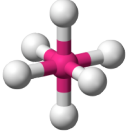
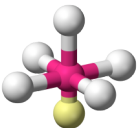
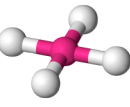
$$\text{Moles of unknown} = \frac{\text{mass of unknown}}{\text{molar mass}} \quad \text{Molar mass} = \frac{\text{mass of unknown}}{\text{moles of unknown}}$$

Molecular Structure (Chapter 9)

- **Molecular Shape:** a molecule's geometry as determined by its bond angles
- **VSEPR** (valence shell electron pair repulsion) model: electron pairs repulse each other and will therefore be as far apart from each other as possible
- **Bonding pair/nonbonding (lone) pair:** a valence electron pair shared between atoms/a nonbonded valence electron pair in the domain of only one atom
- **Electron-domain geometry/molecular geometry:** the geometry of a molecule based on the placement of *all* its electron domains/the geometry of a molecule based on the placement of its *bonded* electron domains only
- **Bond dipole:** a dipole moment due only to the two atoms in a covalent bond
- **Hybridization/hybrid orbitals:** overlapped orbitals that create a new orbital shape
- **Sigma bonds (σ) / Pi (π) bonds**
 - Single bond: 1 sigma
 - Double bond: 1 sigma, 1 pi
 - Triple bond: 1 sigma, 2 pi

How to determine molecular shape:

1. Draw a Lewis Dot Structure
2. Figure out the number of electron domains around the central atom(s)
 - a. A lone pair = one domain
 - b. Any (single/double/triple) bond = one domain
3. The number of electron domains determines **electron-domain** geometry
4. Determine the number of lone pairs
5. The number of lone pairs determines **molecular** geometry

Electron domains	Electron-domain geometry	Molecular Geometry			Bond Angles
		0 lone pair	1 lone pair	2 lone pair	
2	Linear	Linear 	---	---	180°
3	Trigonal planar	Trigonal planar 	Bent 	---	120°
4	Tetrahedral	Tetrahedral 	Trigonal pyramidal 	Bent 	109.5°
5	Trigonal bipyramidal	Trigonal bipyramidal 	See-saw 	T-shaped 	90°, 120°
6	Octahedral	Octahedral 	Square pyramidal 	Square planar 	90°

Practice Problems

- For a molecule with the formula AB_2 the molecular shape is...
 - Linear or bent**
 - T-shaped
 - Linear or T-shaped
 - Trigonal planar
- Which of the following has a square planar molecular geometry?
 - CCl_4
 - XeF_4**
 - XeF_2
 - ICl_3
- The molecular geometry of the H_3O^+ ion is...
 - Tetrahedral
 - Bent
 - Trigonal pyramidal**
 - Octahedral

4. The electron-domain geometry of a sulfur-centered compound is trigonal bipyramidal. The hybridization of the central sulfur atom is...
 - a. sp
 - b. sp²
 - c. sp³
 - d. sp³d**
5. The hybridization of orbitals on the central atom in a molecule is sp. The electron-domain geometry around this central atom is...
 - a. Octahedral
 - b. Linear**
 - c. Trigonal planar
 - d. Tetrahedral
6. The molecular geometry of the BeCl₂ molecule is ... and this molecule is...
 - a. Linear, nonpolar**
 - b. Linear, polar
 - c. Bent, nonpolar
 - d. Bent, polar
7. How many sigma (σ) bonds and pi (π) bonds are there in the O=S–O bond?
 - a. 1 sigma, 1 pi
 - b. 2 sigma, 1 pi**
 - c. 2 sigma, 2 pi
 - d. 3 sigma, 2 pi

WHITE LIGHTNING CHEMISTRY

-**Chemical bond** attraction between atoms

-Types of bonds:

- ionic bonds** (electrostatic forces that hold ions together, e.g., NaCl);
- covalent bonds** (result from sharing electrons between atoms, e.g., Cl₂);
- metallic bonds** (refers to metal nuclei floating in a sea of electrons, e.g., Na).

It's all about *valence electrons*, baby!

- The energy required to separate one mole of a solid ionic compound into gaseous ions is called the **lattice energy**, $\Delta H_{\text{lattice}}$.

$$\text{The strength of any ionic bond} = E = k \frac{Q_1 Q_2}{d}$$

where E is the potential energy of the two interacting charged particles, Q_1 and Q_2 are the charges on the particles, d is the distance between their centers, and k is a constant ($8.99 \times 10^9 \text{ Jm/C}^2$).

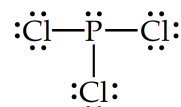
- As Q_1 and Q_2 increase, E increases; and as d increases, E decreases.

-**Polyatomic Ions**

- Covalent, with an overall charge

-**Covalent Bonding**

- The majority of chemical substances do not have the characteristics of ionic compounds.
- A chemical bond formed by sharing a pair of electrons is called a *covalent* bond.
- Both atoms acquire noble-gas electron configurations.



-Multiple Bonds

- One shared pair of electrons is a **single bond** (e.g., H₂).
- Two shared pairs of electrons form a **double bond** (e.g., O₂). $:\ddot{N} \cdot + \cdot \ddot{N} : \longrightarrow :N:::N:$ (or :N≡N:)
- Three shared pairs of electrons form a **triple bond** (e.g., N₂).
- **Bond length** is the distance between the nuclei of the atoms in a bond.
(Distances decrease with the # of bonds)

-Bond Polarity and Electronegativity

- In a **nonpolar covalent bond** the electrons are shared equally.
- In a **polar covalent bond**, one of the atoms exerts a greater attraction for the electrons

-Electronegativity

- The ability of an atom *in a molecule* to attract electrons to itself is its **electronegativity**.
- The electronegativity of an element is related to its ionization energy and electron affinity.
- The Pauling electronegativity scale ranges from 0.7 (Cs) to 4.0 (F).
- Electronegativity increases across a period and decreases down a group.

-Electronegativity and Bond Polarity

- Electronegativity differences close to zero result in nonpolar covalent bonds.
 - The electrons are equally or almost equally shared.
- The greater the difference in electronegativity is between two atoms, the more polar the bond (polar covalent bonds) is.
- There is no sharp distinction between bonding types.

-Dipole Moments

- Molecules like HF have centers of positive and negative charge that do not coincide.
- These are **polar molecules**.
- We indicate the polarity of molecules in two ways:
 - The positive end (or pole) in a polar bond may be represented with a “δ+” and the negative pole with a “δ-”.
- We can also place an arrow over the line representing the bond - the arrow points toward the more electronegative element and shows the shift in electron density toward that atom.
- We can quantify the polarity of the molecule.
 - When charges are separated by a distance, a **dipole** is produced.
 - The **dipole moment** is the quantitative measure of the magnitude of the dipole (μ)
$$\mu = Q r$$
 - when r = the distance between equal and opposite charges Q⁺ and Q⁻
 - The magnitude of the dipole moment is given in *debyes* (D)

Strengths of Covalent Bonds

- The energy required to break a covalent bond is called the **bond enthalpy**, D .
 - That is, for the Cl₂ molecule, $D(\text{Cl-Cl})$ is given by ΔH for the reaction: $\text{Cl}_2(g) \rightarrow 2\text{Cl}(g)$.
- When more than one bond is broken:
$$\text{CH}_4(g) \rightarrow \text{C}(g) + 4\text{H}(g) \quad \Delta H = 1660 \text{ kJ}$$
 - The bond enthalpy is a fraction of ΔH for the atomization reaction:
 $D(\text{C-H}) = \frac{1}{4} \Delta H = \frac{1}{4}(1660 \text{ kJ}) = 415 \text{ kJ}$.
- Bond enthalpy is always a positive quantity.

-Bond Enthalpies and the Enthalpies of Reactions

- We can use bond enthalpies to calculate the enthalpy for a chemical reaction.
- We recognize that in any chemical reaction, bonds need to be broken and then new bonds form.

- The enthalpy of the reaction is given by:
 - the sum of bond enthalpies for bonds broken less the sum of bond enthalpies for bonds formed.
- Where ΔH_{rxn} is the enthalpy for a reaction,

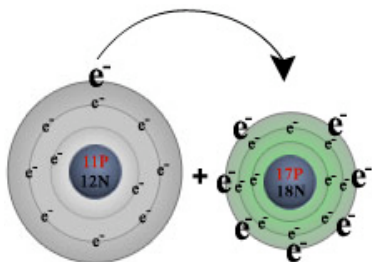
$$\Delta H_{\text{rxn}} = \sum D(\text{bonds broken}) - \sum D(\text{bonds formed})$$
- We illustrate the concept with the reaction between methane, CH_4 , and chlorine:

$$\text{CH}_4(g) + \text{Cl}_2(g) \rightarrow \text{CH}_3\text{Cl}(g) + \text{HCl}(g)$$
 - In this reaction one CH bond and one ClCl bond are broken while one CCl bond and one HCl bond are formed.
 - So $\Delta H_{\text{rxn}} = [D(\text{C}\text{H}) + D(\text{Cl}\text{Cl})] - [D(\text{C}\text{Cl}) + D(\text{H}\text{Cl})] \approx 104 \text{ kJ}$.
 - The overall reaction is exothermic which means that the bonds formed are stronger than the bonds broken.
 - The above result is consistent with Hess's law.

Bond Enthalpy and Bond Length

- The distance between the nuclei of the atoms involved in a bond is called the **bond length**.
- Multiple bonds are shorter and stronger than single bonds.
 - As the number of bonds between atoms increases, the atoms are held closer and more tightly together (multiple bonds are shorter than single bonds)
- Network covalent**: These are giant molecular lattice structures. This implies that strong covalent bonding holds their atoms together in a highly regular, extended network. The bonding between the atoms goes on and on in three dimensions. Because of the large amount of energy needed to break huge numbers of covalent bonds, all **giant covalent network structures** have high melting points and boiling points and are insoluble in water.

Ionic



Metallic

