

Chapter 0: General and Lab Concepts Review

Lesson 0.1 – Atoms, Ions, and Molecules

- *Atom* – smallest unit of an element
- *Molecule* – atom bonded to at least one other atom
- *Ion* – atom that has lost or gained an electron

Lesson 0.2 – Naming Ionic Compounds

- Diatomic elements
 - H₂, N₂, O₂, F₂, Cl₂, Br₂, and I₂
- *Allotropes* – Elemental molecules with different formulas. i.e. O₂ and O₃
- *Ionic compounds* – metal and nonmetal (can be polyatomic too)
- *Molecular compounds* – two or more nonmetals
- Naming ionic compounds
 - 1. Name the metal.
 - 2. Write the metal's oxidation state as a Roman numeral in parenthesis.
 - 3. Name the nonmetal using an "ide" ending
 - ♣ Ex. Iron (II) Chloride, FeCl₂

Lesson 0.3 – Naming Molecular Compounds and Acids

- Naming molecular compounds
 - 1. Give the appropriate numerical prefix to the first element
 - 2. Name the first element using its regular name on the periodic table
 - 3. Give the appropriate numerical prefix to the second element
 - 4. Name the nonmetal using an "ide" ending
 - ♣ Ex. dinitrogen monoxide, N₂O
- Naming binary acids
 - 1. Hydro-
 - 2. Insert the name of element 2, replacing "-ine" with "-ic acid"
 - ♣ Ex. hydrochloric acid, HCl
- Naming oxyacids
 - HClO – hypochlorous acid
 - HClO₂ – chlorous acid
 - HClO₃ – chloric acid
 - HClO₄ – perchloric acid

Lesson 0.4 – Metric Units and Dimensional Analysis

$$\left(\frac{8.41 \times 10^6 \cancel{\text{L}}}{1} \right) \left(\frac{1 \text{ gal}}{3.785 \cancel{\text{L}}} \right) = \text{gal}$$

Lesson 0.5 – Significant Figures and Percent ErrorSignificant Figures

1. Zeroes between nonzero digits are always significant:
 - a. 1005 kg (four sig-figs); 7.03 cm (three sig-figs).
2. Zeroes at the start of a number are never significant:
 - a. 0.02 g (one sig-fig); 0.0026 cm (two sig-figs).
3. Zeroes at the end of a number are significant *only if* the zeroes are AFTER a decimal point:
 - a. 0.0200 g (three sig-figs); 3.0 cm (two sig-figs)

Precision vs. Accuracy

This archer is precise but not accurate.



This archer is precise and accurate.



This archer is imprecise and inaccurate.

Percent Error Calculations

$$\% \text{ error} = \frac{(\text{actual} - \text{theoretical})}{\text{theoretical}} \times 100$$

Lesson 0.6 – Lab Safety

<https://datbootcamp.com/general-chemistry-strategy/general-chemistry-lab-techniques/>

Lesson 0.7 – Weights, Measures, and pH

<https://datbootcamp.com/general-chemistry-strategy/general-chemistry-lab-techniques/>

Lesson 0.8 – Beer-Lambert Law

$$\text{Absorbance} = \epsilon c l$$

- ϵ = molar extinction coefficient (molar absorptivity)
 - Note: Things that are brightly-colored have higher ϵ values
- c = sample's concentration
- l = path length (distance between the light source and the detector in a spectrophotometer)
- **DAT Tip:** If you know the absorbance of a solution, you can calculate its concentration

Chapter 1: Stoichiometry

Lesson 1.1 – The Mole and Molecular Weights

$$1 \text{ mol} = 6.022 \times 10^{23}$$

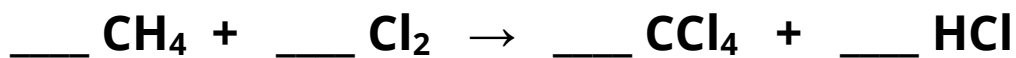
If you have a 150-gram sample of CrO_3 :

- How many moles of CrO_3 do you have?

 - How many oxygen atoms do you have?

 - How many grams of oxygen do you have?
-

Lesson 1.2 – Balancing Reactions



Lesson 1.3 – Molecular Formulas and Percent CompositionEmpirical vs. Molecular Formulas

What is the empirical formula for glucose, $C_6H_{12}O_6$?

Percent Composition

What is the percentage of carbon in $C_6H_{12}O_6$?

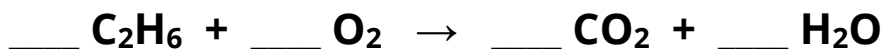
What is the percent composition of nitrogen in morphine, whose molecular formula is $C_{17}H_{19}NO_3$?

What is the empirical formula of a compound that is 17.3% H and 82.7% C?

If your compound's actual molecular weight is 58.123 amu, then what is its molecular formula?

Lesson 1.4 – Product and Reactant Amounts

How many moles of C_2H_6 are needed to react with excess O_2 to form 22 moles of CO_2 in the following reaction?



How many grams of C_2H_6 does this represent?

Lesson 1.5 – Limiting Reactants

If you react 6 grams of C_2H_6 (M.W. = 30 g/mol) with 9.6 grams of O_2 (M.W. = 32 g/mol), which reactant will run out first?

Under these conditions, how many grams of H_2O will this reaction produce?

Lesson 1.6 – Theoretical and Percent Yields

Suppose you react 81 grams of Al (atomic mass = 27 g/mol) with 213 grams of Cl_2 (M.W. = 71 g/mol), and you obtain 133.5 grams of $AlCl_3$, according to the following equation:



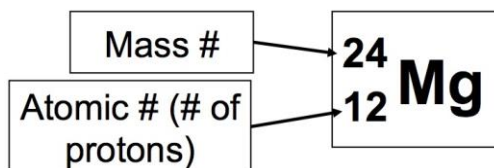
Please calculate:

1. The limiting reactant?
2. The theoretical yield?
3. The percent yield?
4. How many grams of the excess reagent are left over, once the reaction is completed?

Chapter 2: Atomic and Electronic Structure

Lesson 2.1 – Atomic Structure and the Bohr Model

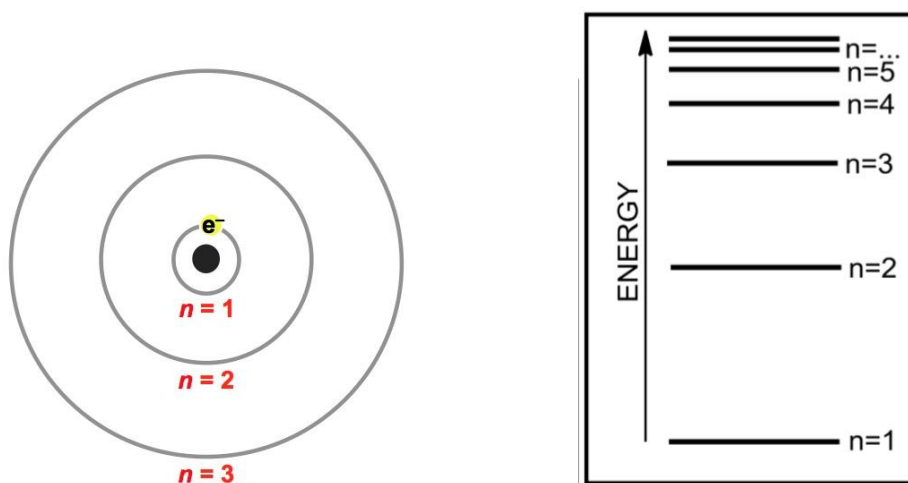
Atomic Structure



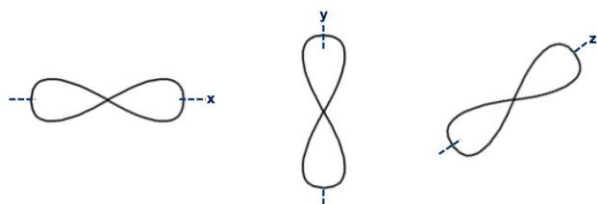
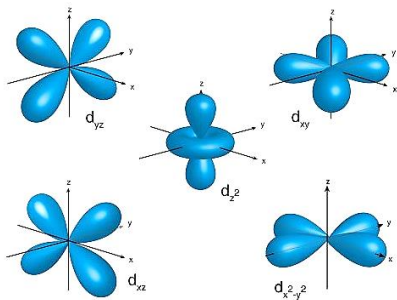
- Mass Number = (Protons) + (Neutrons)
- Neutrons = (Mass Number) - (Protons)

Bohr Model of the Atom

- Electrons distance from their nuclei are **quantized**

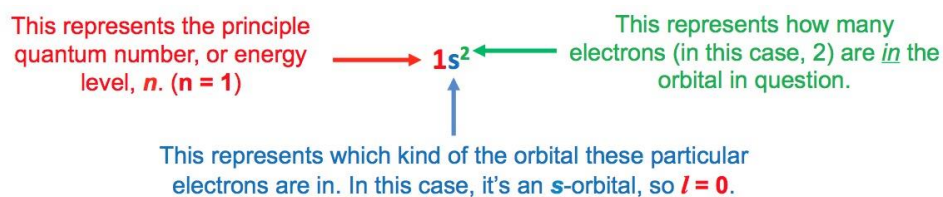


The distance between each energy shell and the *next* shell above it gets smaller as you get further away from the nucleus.

Lesson 2.2 – Atomic OrbitalsS-orbitalsP-orbitalsD-orbitals**Lesson 2.3 – Quantum Numbers**

#	Name	What	Range
n	principal	Shell (distance from nucleus)	[1...infinity]
l	azimuthal	subshell (type of orbital)	[0...(n-1)] l = 0 (s) l = 1 (p) l = 2 (d) l = 3 (f)
m_l	magnetic	specific orbital (orientation in space)	[-l...+l]
m_s	spin	Up or down	+1/2 or -1/2

Lesson 2.4 – Electron Configuration



- What is the electron configuration of oxygen?
- Cool video on stacking of orbitals and how atoms really look:
<http://www.youtube.com/watch?v=sMt5Dcex0kg>

Lesson 2.5 – Condensed Electron Configuration, Valence, and Energy Diagrams

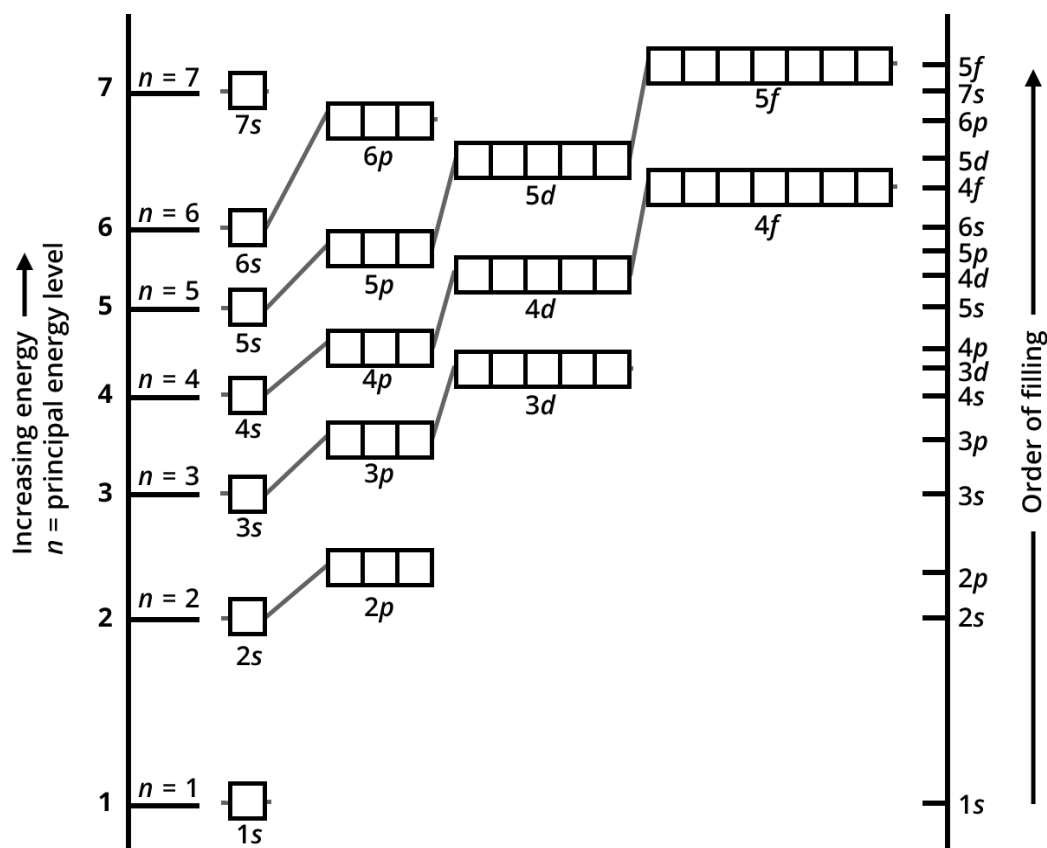
Condensed Electron Configuration

- What is the condensed electron configuration of bromine?

Valence Electrons

- What are bromine's valence electrons?
- How many valence electrons does titanium have?
- When does the d-block count toward an atom's number of valence electrons?

Continue to next page...

Lesson 2.5 – Condensed Electron Configuration, Valence, and Energy Diagrams (Continued)Energy Diagrams

- *Aufbau Principle* – Electrons fill the lowest energy orbitals first
- *Hund's Rule* – Don't pair up electrons until you have to.
- *Pauli Exclusion Principle* – no two electrons in the same atom can have the same four **quantum numbers**. In other words, no two electrons in the same atom can have the exact same address.

Lesson 2.6 – Electron Configuration Exceptions (Cr and Cu)

- What are the 5 exceptions you need to know, and what are their electron configurations?

Lesson 2.7 – Excited Electron Configurations

- Electrons can absorb a photon and be promoted to a higher-energy shell or orbital.

Lesson 2.8 – Paramagnetic vs. Diamagnetic

When I hear ...	I think ...	I then think ...
Paramagnetic	"Unpaired"-a-magnetic	Attracted to magnets because it has unpaired electrons.
Diamagnetic	The other one (all paired)	Slightly repelled by magnets because it has all paired electrons.

- If an element has an **ODD** number of electrons, then it's **paramagnetic**.
- If it has an **EVEN** number of electrons, then it can be **either** *paramagnetic* or *diamagnetic*; you have to fill out the electron configuration energy diagram to find out.
- Hint:** Liquid oxygen is paramagnetic, liquid nitrogen is diamagnetic.

Lesson 2.9 – Emission Spectra, Heisenberg Uncertainty, Photoelectric Effect

$$E_{\text{photon}} = hf = \frac{hc}{\lambda}$$

↑ Energy ≈ ↑ frequency ≈ ↓ λ

f = the photon's **frequency** (this can be different for different photons)

c = **speed of light**, which is $3.0 \times 10^8 \text{ m/s}$

h = **Planck's constant**, which is $6.63 \times 10^{-34} \text{ J} \cdot \text{s}$

λ = the photon's **wavelength**

Heisenberg Uncertainty

- It is impossible to determine a subatomic particle's **position** and its **momentum** with perfect accuracy.

Photoelectric Effect

$$\text{Kinetic Energy}_{\text{electron}} = E_{\text{photon}} - \Phi$$

Φ = work function (the minimum amount of energy required to ionize the electron)

- In order to expel an energized electron, the **Kinetic Energy** must be greater than zero.





Chapter 3: Molecular Structure and Geometry

Lesson 3.1 – Types of Bonds and Compounds

Types of Bonds

- *Ionic bonds* – transfer electrons between elements with large difference in electronegativity (usually metal with nonmetal, although polyatomic ions can be made of all nonmetals)
- *Covalent bonds* – sharing of electrons between two elements with similar electronegativity (between two nonmetals).

Table of Bonds

Type	Properties	What keeps them together?	Examples
 Ionic	high melting points, high boiling points, brittle, hard	Ionic interactions (lattice energy)	NaCl, MgO
Molecular	low melting points, do not conduct electricity	Intermolecular Forces	
 Network Covalent	high melting points, high boiling points, hard, do not conduct electricity	Network of covalent bonds	C (diamond, graphite) SiO₂ (quartz)
 Metallic	variable hardness and melting points, conduct electricity, conduct heat, lustrous (shiny), malleable, ductile	Metallic Bonding	Fe, Mg

Polyatomic Ions

- <https://datbootcamp.com/blog/common-ions-you-need-to-know>

Octet Rules

- Every atom wants to feel like it has 8 electrons
- **Exceptions:** Hydrogen only wants 2 e⁻, Beryllium only wants 4 e⁻, Boron and Aluminum want 6 e⁻ usually
- Third row and lower can have more than 8 electrons if needed

Lesson 3.2 – Lattice Energy

What is Lattice Energy?

- *Lattice energy* is the energy required to completely separate an ionic compound's *cations* from its *anions*.
 - $\text{NaCl}_{(s)} \rightarrow \text{Na}^+_{(g)} + \text{Cl}^-_{(g)}$
 - **Larger charges** = larger lattice energy
 - **Shorter bond distance (smaller ions)** = larger lattice energy
-

Lesson 3.3 – Lewis Structures

Octet Rules

- Every atom wants to feel like it has 8 electrons
- **Exceptions:** Hydrogen only wants 2 e⁻, Beryllium only wants 4 e⁻, Boron and Aluminum usually want 6 e⁻
- Third row and lower can have more than 8 electrons if needed

How to Draw Lewis Structures

1. Count up the valence electrons for every atom in your formula. Add one electron to your count for each negative charge. Subtract one electron from your count for each positive charge.
2. Write down the symbols for all the atoms in your molecule. Usually, the least electronegative atom goes in the center. Connect each atom with a single bond, remembering that each single bond represents two electrons.
3. Use your remaining electrons to complete everyone's octet by adding *lone-pair electrons* (also called *nonbonding pairs*) around each atom, leaving the central atom for last.
4. Place any leftover electrons on the central atom, even if doing so violates the octet rule.
5. If there are not enough electrons to give the central atom an octet, form multiple bonds (like double or triple bonds) to complete the central atom's octet.

Try it on your own! (Answer in Solution Video 3.3)

1. Which of the following molecules contains a triple bond?

- I. N_2
 - II. CH_3CN
 - III. C_2H_2
- A. I only
 - B. I and III only
 - C. II and III only
 - D. III only
 - E. I, II, and III

Lesson 3.4 – Formal Charge and Resonance ContributorHow to Count Bonds

- Each single bond in a molecule is one σ (*sigma*) bond.
- All double bonds contain one σ and one π (pi) bond.
- Triple bonds contain one σ and two π bonds

Formula Charges

Formal Charge = (valence electrons) – (dots) – (bonds)
--

- Major resonance contributor usually the one with the least amount of formal charges.

Lesson 3.5 – Bond Order

- **Bond order** is the average length of each bond, when multiple **contributors** are possible.
- Count up each individual bond and divide by the total number of ‘things’ around the central atom.

Lesson 3.6 – Hybridization and Molecular Geometry

Electron Domains ('things')	Hybridization	Bond Angles	Electron Domain Geometry	Non-bonding Pairs of e ⁻	Molecular Geometry
2	sp	180°	Linear	0	Linear
3	sp ²	120°	Trigonal planar	0	Trigonal planar
				1	Bent
4	sp ³	109.5°	Tetrahedral	0	Tetrahedral
				1	Trigonal pyramidal
				2	Bent
5	sp ³ d	90°, 120°	Trigonal bipyramidal	0	Trigonal bipyramidal
				1	See-saw
				2	T-shaped
				3	Linear
6	sp ³ d ²	90°	Octahedral	0	Octahedral
				1	Square pyramidal
				2	Square planar

Lesson 3.7 – PolarityHow to Determine Polarity

1. Draw your molecule's *Lewis structure* flat on your paper and then spread out all of its groups (including its lone pairs) as far as possible, around its central atom.
2. Draw arrows between every atom in the molecule, going from the less electronegative atom (A) to the more electronegative atom (B) in each bond.
3. Answer the following question: "If my central atom were a truck stuck in the mud being pulled in the directions indicated by the arrows, would the truck move?" If so, then your molecule is *polar*. If not, then it's *nonpolar*.

Try it on your own! (Answer in Solution Video 3.7)

1. Which of the following molecules would be most soluble in toluene? Which one would be the least?
 - A. $\text{CH}_3\text{CH}_2\text{CH}_2\text{OH}$
 - B. $\text{CH}_3\text{CH}_2\text{OH}$
 - C. $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{OH}$
 - D. CH_3OH
2. Predict whether each of the following molecules is *polar* or *nonpolar*:
 - a) CCl_4
 - b) NH_3
 - c) SF_4
 - d) XeF_4

Chapter 4: Periodic Trends

Lesson 4.1 – Periodic Table and Descriptive Chemistry

Periodic Table

- Periods are rows on the periodic table
- Groups are vertical columns on the periodic table
- Elements in the same group tend to have similar chemical properties

Important Groups to Know

- *Alkali metals* – low ionization energies, react explosively with water, form ionic compounds
- *Alkaline earth metals* – low ionization energies, don't react as violently with water
- *Halogens* – highly electronegative, high electron affinities, highly reactive with metals, good oxidizing agents (they want to be reduced)
- *Noble gases* – inert (unreactive) gases
- *Transition metals* – form colorful compounds due to the way their d-orbitals behave
- *Oxygen group* – O₂ is a good oxidizing agent, but O₃ is even better oxidizing agent.
 - Likes to react with metals to form metal oxides.

Lesson 4.2 – Atomic Size (Radius) and Bond Lengths

Atomic Size (Radius)

- Atoms get LARGER as you go DOWN and to the LEFT (**Francium** is largest element)
- Why do atoms get larger as you go DOWN a group?
 - Higher energy “n” shell = larger size
- Why do atoms get smaller as you go ACROSS a period, from left to right?
 - ↑ Effective nuclear charge (Z_{eff}) = ↓ size

Bond Length

- The distance between the nuclei of two atoms

$$\text{Bond length} = (\text{atomic radius}_1 + \text{atomic radius}_2)$$

Lesson 4.3 – Isoelectronic Series Sizes

Ion Sizes

- More electrons (anions) = \uparrow size
- Less electrons (cations) = \downarrow size

Which of the following atoms would be the largest? Which one would be the smallest?

- A. O
- B. O⁻
- C. O²⁻
- D. O⁺

Isoelectronic Series

- Group of elements that all have the same number of electrons
- More protons = more attraction to electrons (Z_{eff}) = smaller size atom

In the following isoelectronic series, which one would be the largest? The smallest?

- A. O²⁻
- B. F¹⁻
- C. Ne
- D. Na¹⁺
- E. Mg²⁺

Lesson 4.4 – Ionization Energy

Ionization Energy

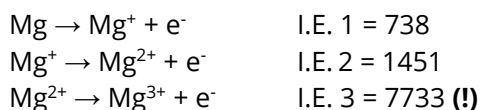
- *Definition* – Energy required to remove an electron from an atom
- *Trend* – Increases as you go UP and to the RIGHT on periodic table (**Helium** has highest I.E.)

Exceptions

- Beryllium has a full 2s² subshell, and Nitrogen has half full 2p³ subshell. Takes more energy to strip these electrons since they are relatively stable.
- Therefore, Beryllium has a higher ionization energy than Boron, and Nitrogen has a higher ionization energy than Oxygen.

Successive Ionization Energies

- Once you remove all the valence electrons, ionization energy increases a lot



Lesson 4.5 – Electronegativity and Electron Affinity

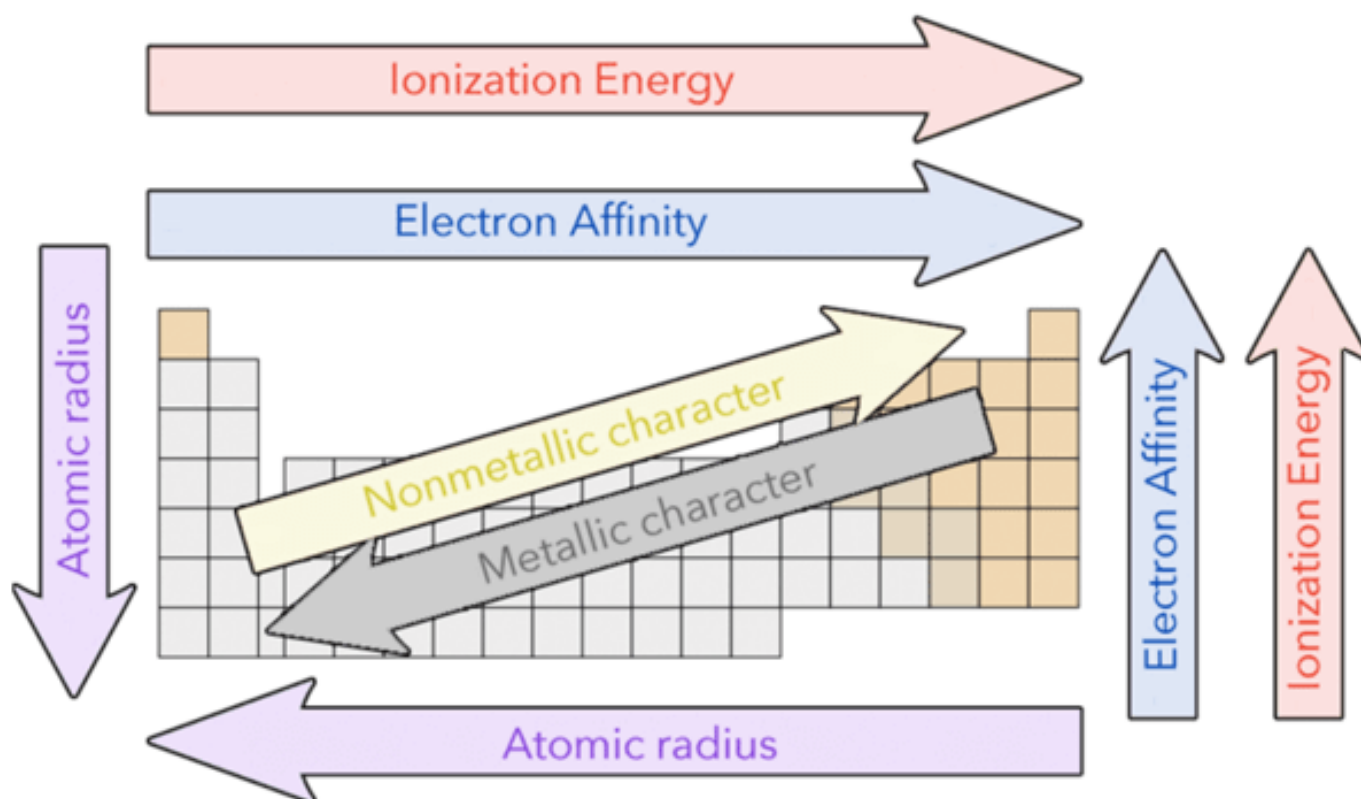
Electronegativity

- *Definition* – How 'thirsty' an atom is for electrons
- *Trend* – Increases as you go UP and to the RIGHT (**Fluorine** is most electronegative)
 - *Note*: Noble gases are excluded from this trend, they have a full octet already

Electron Affinity

- *Definition* – the energy **given off** when an atom gains an electron
 - *Tip*: Electron affinity is an **exothermic** reaction
- *Trend* – Increases as you go UP and to the RIGHT (**Chlorine** has highest electron affinity)

Summary of Periodic Trends



Chapter 5: Gases

Lesson 5.1 – Temperature, Volume, and Pressure (Gases)

Temperature

$$K = ^\circ C + 273$$

Volume

$$\text{Liters, } 1\text{cm}^3 = 1\text{mL} = 1\text{cc}$$

Pressure

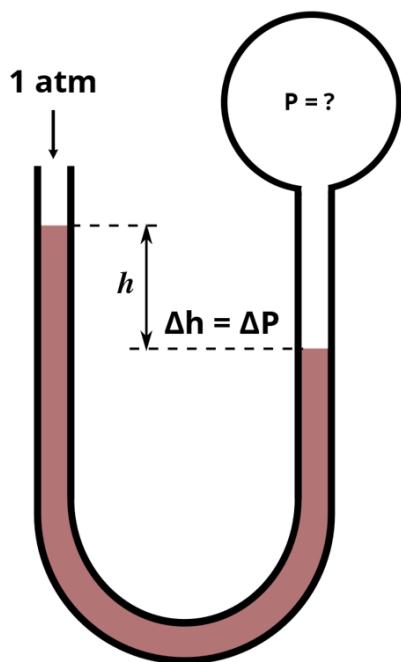
Pressure is caused by gas molecules hitting the side of a container

$$P = F/A$$

$$1\text{ atm} = 760\text{ torr} = 760\text{ mmHg}$$

Lesson 5.2 – How to Measure Gas Pressure in a Lab

Mercury Manometer



The unknown gas pushes down the tube of mercury by 40mm. What is the pressure of the gas, assuming we are at sea level?

Lesson 5.3 – Kinetic Molecular Theory and Ideal Gas AssumptionsIdeal Gas Assumptions

- The volume or size of each individual gas molecule is insignificant.
 - Most accurate at low pressures
 - Gas molecules' collisions with each other are perfectly elastic. No intermolecular forces.
 - Most accurate at high temperatures, and low IM forces
 - The average kinetic energy of a gas depends only on the system's temperature.
- Gases behave most ideally under LOW pressures and HIGH temperatures.

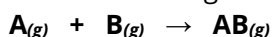
Lesson 5.4 – Combined Gas Law, Ideal Gas LawGas Laws

Boyle's Law	$V \propto \frac{1}{P}$	1 mol of gas = 22.4 L at STP R = 0.0821 L•atm/mol•K STP = 1 atm and 273K
Charles' Law	$V \propto T$	
Avogadro's Law	$V \propto n$	
Combined Gas Law	$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$	
Ideal Gas Law	$PV = nRT$	

1. A flask contains 2.0 mol of He gas at 25°C and 1.00 atm. How much He gas, in grams, must be added to increase the pressure to 2.00 atm at constant temperature and volume?

2. A 0.02 mol sample of an ideal gas in a sealed 1.0 L container was heated from 27 °C to 227°C. What is the pressure of the gas at this new temperature?

3. Consider the gaseous reaction:



A mixture of 0.04 mole of gas A and 0.1 mole of gas B was allowed to react at 300 K in a 1 L sealed container. At the end of the reaction, what is the total pressure of the mixture? (The temperature and volume remain constant during the reaction.)

Lesson 5.5 – Gas Density, Dalton’s Law of Partial Pressures, Graham’s Law of EffusionGas Density

$$\text{Density} = \frac{P(\text{MM})}{RT} = \frac{m}{v} \quad \text{MM} = \text{Molar mass}$$

Dalton’s Law of Partial Pressures

$$P_{\text{total}} = P_A + P_B + \dots \quad P_A = \chi_A P_{\text{total}} \quad \chi_A = \text{mol fraction}$$

Graham’s Law of Effusion

$$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$$

Lesson 5.6 – Real Gas Equation

$$\left(P + \frac{an^2}{V^2}\right)(V - nb) = nRT$$

$\frac{an^2}{V^2}$ corrects for intermolecular forces

$-nb$ corrects for volume

Chapter 6: Liquids and Solids

Lesson 6.1 – Types of Intermolecular Forces

Intermolecular Forces

Ion-dipole – Intermolecular force between ions and polar substances (NaCl in water)

Hydrogen Bonding – N-H, F-H, O-H bond in a pure substance.

Dipole-dipole – Intermolecular force between polar molecules

London Dispersion Forces (Van der Waals) – Intermolecular force between all molecules

– Temporary dipole, includes polar and nonpolar molecules

– The greater the weight, the greater the London dispersion forces.

Lesson 6.2 – Effects of Intermolecular Forces

Higher Intermolecular Forces Leads to:

Higher Boiling Point

Higher Heat of Vaporization

Higher Viscosity

Higher Surface Tension

LOWER Vapor Pressure

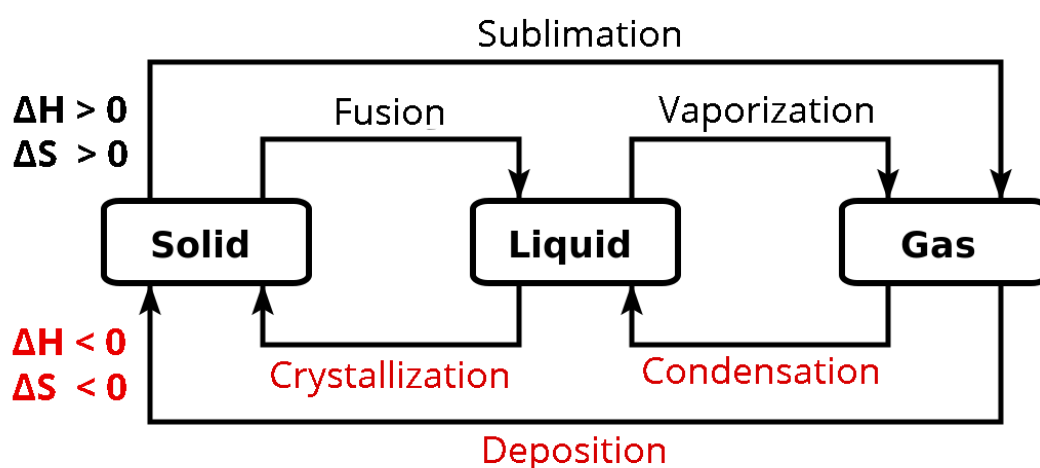
Lesson 6.3 – Structures of Solids

Types of Solids

Type of Solid	Interaction	Properties	Examples
Metallic	Metallic Bonding	Variable hardness and melting point, conductive	Fe, Mg
Ionic	Ionic bonding	High melting point, brittle, hard	NaCl, MgO
Covalent Network	Network of Covalent bonds	High melting point, hard, non-conductive	C (diamond, graphite), SiO ₂ (quartz)
Molecular	H-bonding, dipole-dipole, London dispersion	Low melting point, soft, non-conductive	H ₂ O, CO ₂

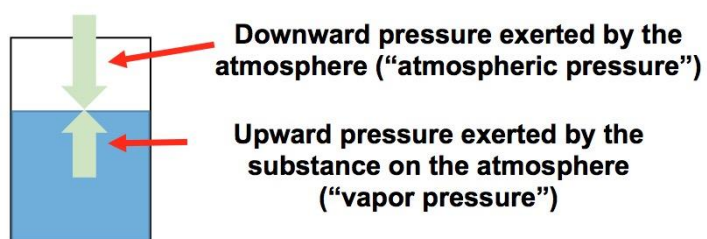
Lesson 6.3 – Structures of Solids (Continued)Cubic Unit Cells

1. Simple Cubic – 1 atom per unit cell. Atoms at corners of unit cell.
2. Body-Centered Cubic – 2 atoms per unit cell. Atoms at corners and body center of unit cell.
3. Face-Centered Cubic – 4 atoms per unit cell. Atoms at corners and face centers of unit cell.

Lesson 6.4 – Phase Changes**Lesson 6.5 – Vapor Pressure and Boiling Point**

Boiling Point – When the vapor pressure = external pressure

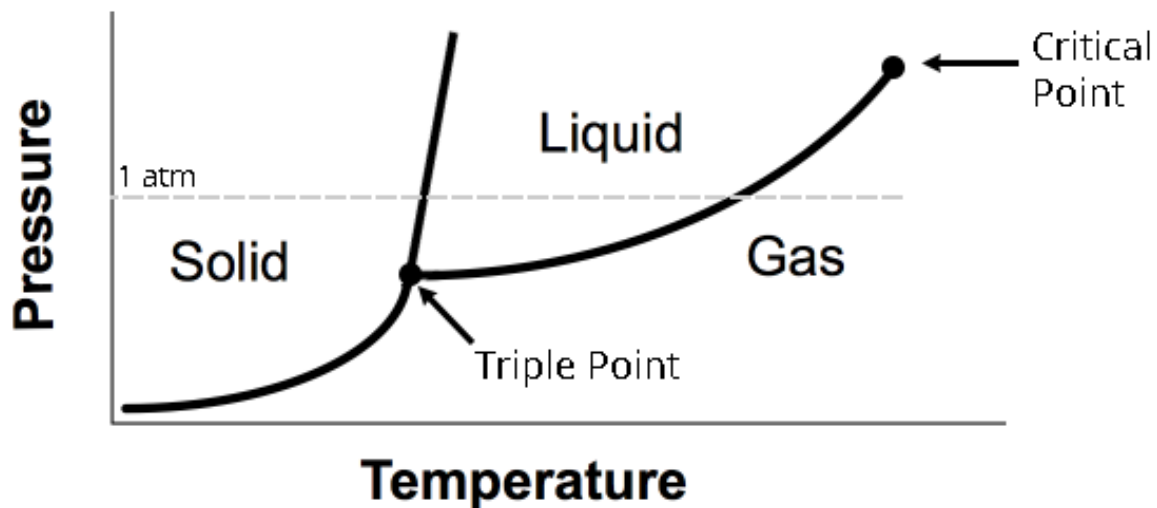
\uparrow boiling point \approx \downarrow vapor pressure



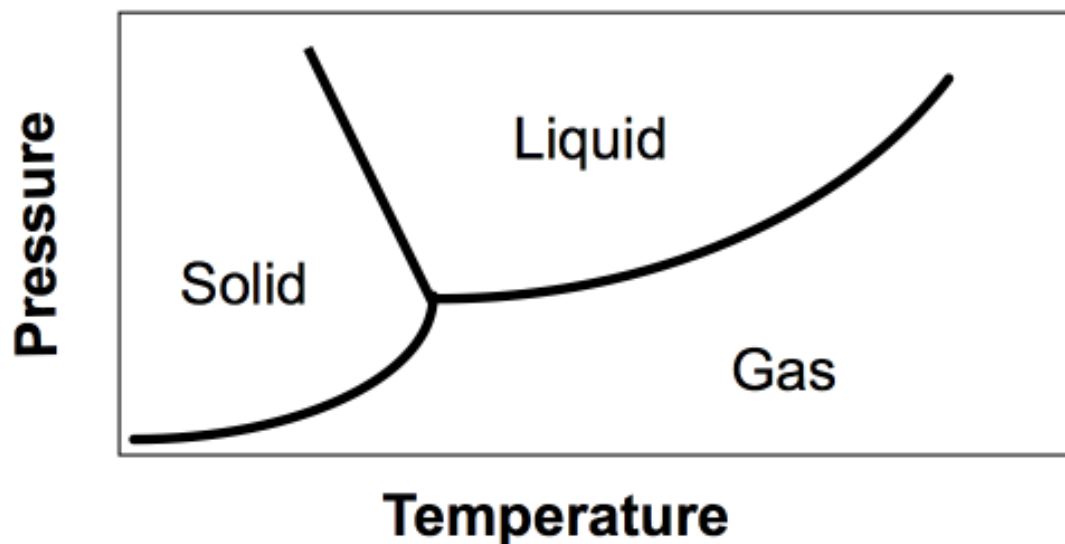
Lesson 6.6 – Phase Diagrams

Identify:

- Lines of equilibrium
- Triple point
- Critical point
- Normal melting point and boiling point

**Lesson 6.7 – Phase Diagram of Water**

- Water has a negative solid-liquid line of equilibrium slope



Chapter 7: Chemical Solutions

Lesson 7.1 – Solubility in Polar and Nonpolar Compounds

How to Determine Polarity

1. Draw your molecule's Lewis structure flat on your paper and then spread out all of its groups (including its lone pairs) as far as possible, around its central atom.
2. Draw arrows between every atom in the molecule, going from the less electronegative atom (**A**) to the more electronegative atom (**B**) in each bond, like this:
3. Answer the following question: "If my central atom were a truck stuck in the mud being pulled in the directions indicated by the arrows, would the truck move?"

"Like Dissolves Like"

Examples

Which of the following, when added to carbon tetrachloride, will form a miscible solution?

- I. hexane
- II. water
- III. propane
- IV. ethyl acetate

Arrange the following alcohols in order of INCREASING solubility in water.

- I. methanol, CH_3OH
- II. ethanol, $\text{CH}_3\text{CH}_2\text{OH}$
- III. n-propanol, $\text{CH}_3\text{CH}_2\text{CH}_2\text{OH}$

Lesson 7.2 – How to Calculate Concentration

Vocabulary

Solution – a uniform mixture of two or more substances

Solvent – the substance in a solution that is present in a larger amount

Solute – the substance in a solution that is present in a smaller amount

Unsaturated – less than the maximum amount of a solute is dissolved in a solvent

Saturated – the maximum amount of a solute is dissolved in a solvent

Concentration

$$\text{Molarity (M)} = \frac{\text{moles}_{\text{solute}}}{L_{\text{solution}}}$$

$$\text{Molality (m)} = \frac{\text{moles}_{\text{solute}}}{\text{kg}_{\text{solvent}}}$$

Lesson 7.2 – How to Calculate Concentration (Continued)Examples

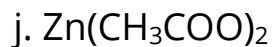
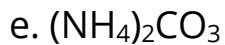
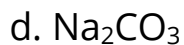
You have 2.0 L of 0.5 M NaCl. How many grams of NaCl do you have?

What mass (in grams) of MnSO_4 (150 g/mol) is required to prepare a 5.0 *m* manganese sulfate solution with 500 g of water?

Lesson 7.3 – Solubility RulesSolubility of Ionic Compounds in Water

1. Most Group 1 metal cations, NO_3^- , ClO_4^- , $\text{C}_2\text{H}_3\text{O}_2^-$, and NH_4^+ salts are **soluble**.
2. Most Ag^+ , Pb^{2+} , S^{2-} , OH^- , Hg_2^{2+} , CO_3^{2-} , and PO_4^{3-} salts are **insoluble**.
3. The solubles generally trump the insolubles.

Predict if the following are soluble or insoluble in water:

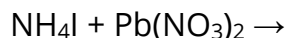


Lesson 7.4 – Net Ionic Equations

To generate a net ionic equation:

1. Balance the chemical equation.
2. Use your solubility-rule knowledge to label everything as (s), (l), (g) or (aq).
3. Cut all the (aq)'s "in half," separating their cations from their anions.
4. Cancel out all the species that are exactly the same on both sides. The canceled-out species are called spectator ions.

Example



- Predict the products
- Total ionic equation
- Net ionic equation
- Spectator ions

Lesson 7.5 – Solubility of Gases and Solids, Henry's Law

Solubility of Solids and Gases

1. Solids are more soluble at higher temperatures (melting sugar in hot coffee)
2. Gases are less soluble at higher temperatures (more energy helps them escape the liquid)
3. Gases are more soluble at higher pressures (more pressure forces gas into the liquid)

Henry's Law

$$P_A = k_H[A]$$

Example

Nitrogen gas' solubility at ambient temperature and 1 atm of pressure is 6.8×10^{-4} mol/L. If the pressure of nitrogen gas in the air is reduced to 0.5 atm, then what is its new concentration?

Lesson 7.6 – Colligative PropertiesFreezing point depression

$$\Delta T_F = -iK_F m$$

Boiling point elevation

$$\Delta T_B = iK_B m$$

Examples

Which of the following solutes would decrease water's freezing point by the greatest amount, per mole of solute added?

- A. NaCl
- B. NH_4NO_3
- C. NH_4Cl
- D. Na_2SO_4
- E. Sucrose

Which of the following solutes would change water's boiling point by the greatest amount, per mole of solute added?

- A. 1.2 *m* $\text{C}_6\text{H}_{12}\text{O}_6$
- B. 0.8 *m* $\text{Ca}(\text{NO}_3)_2$
- C. 1.0 *m* KCl

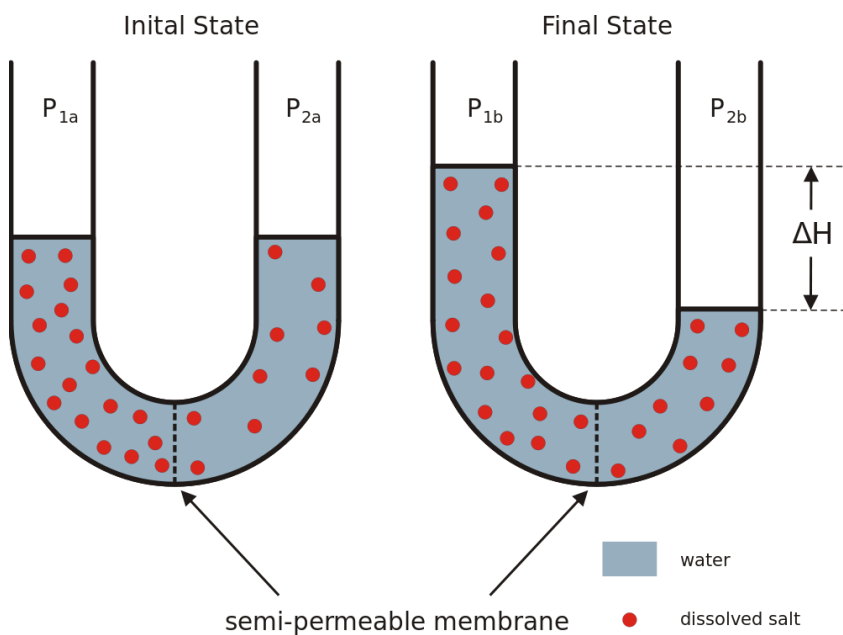
If you have a 1.0 *m* aqueous solution of NaCl, by how much will it increase the water's boiling point, if $K_B = 0.512$ C/*m*? In other words, what is the boiling point elevation (increase)?

Lesson 7.7 – Raoult's LawVapor pressure depression

$$P_A = \chi_A P_{\text{pure}}$$

Lesson 7.8 – Osmosis

$$\Pi = iMRT$$

**Example**

Arrange the following aqueous solutions in order of decreasing osmotic pressure.

- I. 0.20 M KCl
- II. 0.30 M urea
- III. 0.50 M sucrose

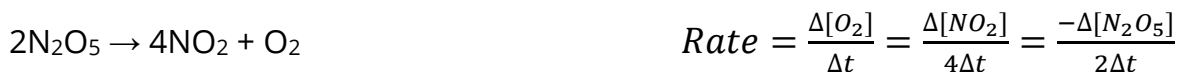
Chapter 8: Chemical Kinetics

Lesson 8.1 – Relative Rate Expressions

Chemical Kinetics – is the study of how fast reactions occur, or reaction rate

Thermodynamics – tells us if a reaction will even happen, but not how fast it'll happen

Rate Expressions



If O_2 is being produced at a rate of 2 M/s, how fast is NO_2 being produced?

If O_2 is being produced at a rate of 2 M/s, what is the rate of change of N_2O_5 ?

Lesson 8.2 – Rate Laws and Constants

Rate Laws

Overall reaction: $\text{A} + \text{B} \rightarrow \text{C} + \text{D}$ rate = $k[\text{A}]^m[\text{B}]^n$

- Products do not appear in the rate law
- k is the rate constant
- 'm' and 'n' have to be determined experimentally. They are not the coefficients in the balanced rxn.
- 'm' and 'n' are the reaction orders. A reaction's overall reaction order is $m+n$.

Rate Constant

0 order units of $k = \text{M}^1 \cdot \text{s}^{-1}$

1st order units of $k = \text{s}^{-1}$

2nd order units of $k = \text{M}^{-1} \cdot \text{s}^{-1}$

3rd order units of $k = \text{M}^{-2} \cdot \text{s}^{-1}$

Lesson 8.2 – Rate Laws and Constants (Continued)Overall Reaction: $2A + B \rightarrow C$

	$[A]_{\text{initial}}$ (M)	$[B]_{\text{initial}}$ (M)	Initial rate (M/s)
Trial 1	0.10	0.3	0.01
Trial 2	0.20	0.3	0.04
Trial 3	0.10	0.6	0.02

What is the rate law for this reaction?

What is the overall reaction order?

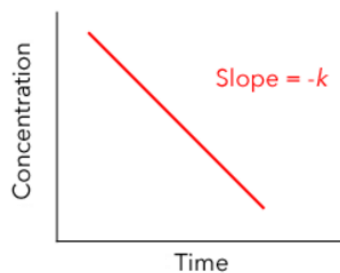
What is the reaction's rate constant (units)?

Lesson 8.3 – Integrated Rate LawsUnits for y-axis of integrated rate law

zero-order = [concentration]

first-order = \ln [concentration]second order = $1 /$ [concentration]Integrated rate law graphs

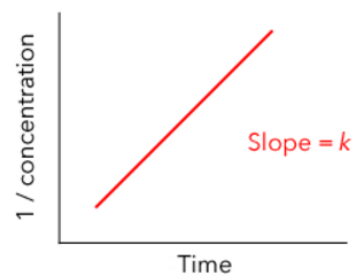
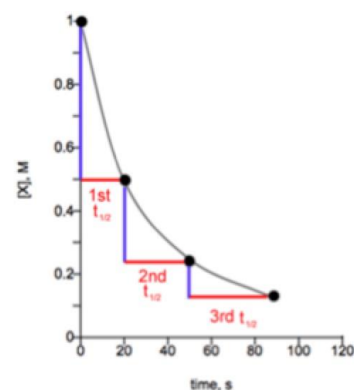
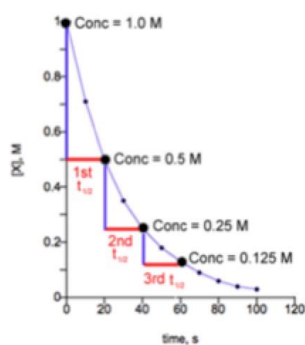
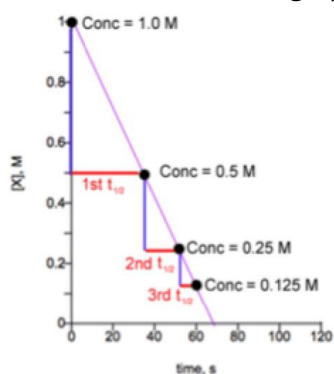
If the reaction is zero order...



If the reaction is first order...



If the reaction is second order...

Concentration [X] vs. time graphszero-order reactions:

- The plot is a straight line
- \downarrow Half-life \approx \downarrow concentration

first-order reactions:

- NOT a straight line
- Half-life constant is independent of concentration

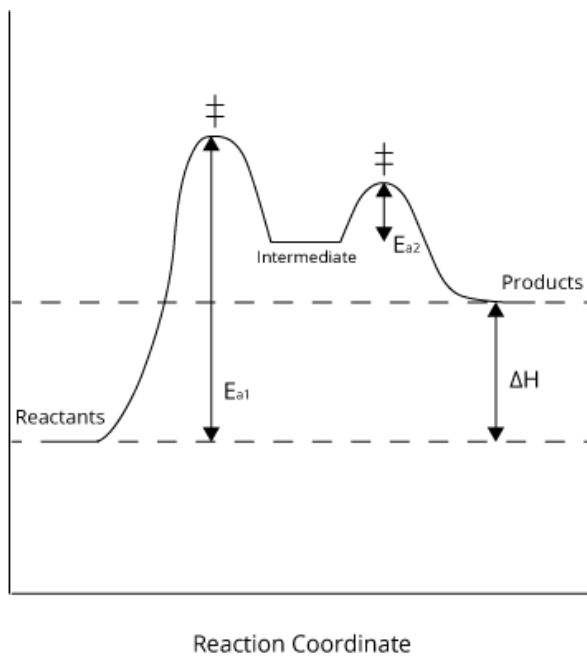
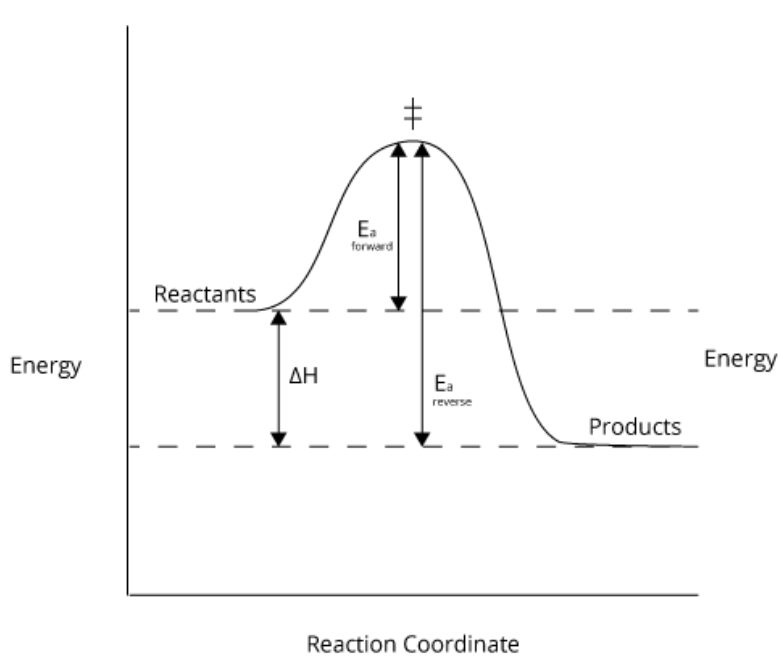
second-order reactions:

- NOT a straight line
- \uparrow Half-life \approx \downarrow concentration

Lesson 8.4 – Reaction Coordinate Diagrams

Identify in each:

1. ΔH (exothermic or endothermic)
2. E_a
3. Transition states (activated complex)
4. Intermediates
5. Rate-determining step (slow step)

**Lesson 8.5 – Catalysts**

1. Speed up a reaction without being consumed in the process
2. Works by lowering the activation energy
3. Provides an alternate pathway (different mechanism) between reactants and products
4. Does NOT shift equilibrium, you just reach equilibrium faster.

Lesson 8.6 – Collision Theory and Arrhenius EquationCollision Theory

1. Both molecules must collide*
2. Both must collide with enough energy*
3. Both must collide in the correct 3D orientation

*Increasing the temperature increases collision frequency and percentage of high-energy collisions.

Arrhenius Equation

$$k = Ae^{-E_a/RT}$$

$\uparrow T = \uparrow k = \uparrow \text{rate}$

$\downarrow E_a = \uparrow k = \uparrow \text{rate}$

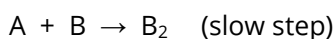
Lesson 8.7 – Reaction Mechanisms

Intermediates – are produced in the middle of the reaction and get used up before the end.

Rate law – rate = $k \times$ (concentration of each reactant in the slow step, multiplied by each other).

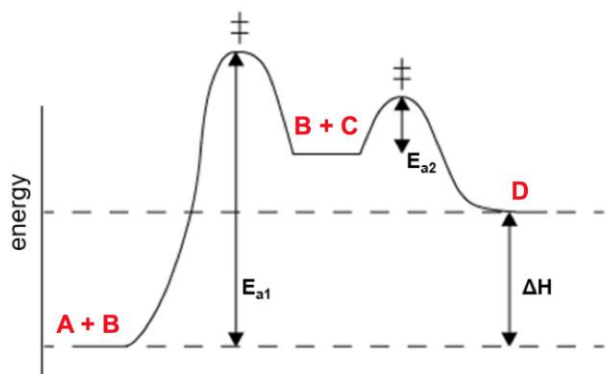
Overall reaction – add up the elementary steps and cancel out intermediates.

A certain chemical reaction has the following elementary steps:



1. What is the intermediate?
2. What is the overall rate law?
3. What is the overall reaction?

1. What are the elementary steps?
2. What is/are the intermediate(s)?
3. What is the overall rate law?
4. What is the overall reaction?



Chapter 9: Chemical Equilibria

Lesson 9.1 – Dynamic Equilibrium and Equilibrium Constants

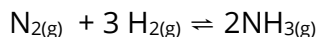
Dynamic equilibrium: Forward Rate = Reverse Rate

Equilibrium Constant

$$K_c = \frac{[\text{products}]}{[\text{reactants}]} \quad K_p = \frac{P_{\text{products}}}{P_{\text{reactants}}} \quad K_{eq} = \frac{k_{\text{forward}}}{k_{\text{reverse}}}$$

- Only aqueous or gaseous are included in equilibrium expression
- Only a change in TEMPERATURE can change equilibrium constant k

Keq	Meaning
K more than 1	Products favored at eq.
K less than 1	Reactants favored at eq.



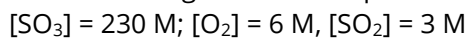
1. What is the rate constant expression in terms of concentration, K_c ?
2. What is the rate constant expression in terms of pressure, K_p ?
3. If $K_c = 10$, then what is K_c for the reverse reaction, $2\text{NH}_3 \rightleftharpoons \text{N}_2 + 2 \text{H}_2$?

Lesson 9.2 – Reaction Quotient (Q vs. K)



Which side of the reaction is favored?

Is the following mixture at equilibrium?

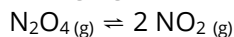


Which direction does it need to move to achieve equilibrium?

Q	Result
Q is less than K	shift right (products)
Q is greater than K	shift left (reactants)
Q = K	at equilibrium

Lesson 9.3 – Le Chatelier’s Principle

Le Chatelier’s Principle – If a system at equilibrium is disturbed, then it will shift in whichever direction it has to, to restore equilibrium.

Changing Concentrations (Only gaseous/aqueous compounds!)

In which direction will the reaction shift to restore equilibrium if:

- N_2O_4 is removed?
- NO_2 is removed?
- N_2O_4 is added?
- NO_2 is added?

Changing Temperature (Exothermic/Endothermic)

In which direction will the reaction shift to restore equilibrium, if we remove or add heat?

Changing Volume and/or Pressure

In which direction will the equilibrium shift when:

- If we remove or add heat?
- The pressure is increased by adding $\text{N}_2(g)$?
- The volume is increased?
- The volume is decreased?
- The pressure is decreased?

Try it yourself!

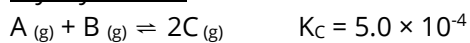
In which direction will the equilibrium shift when:

- The temperature is increased?
- The temperature is decreased?
- Pressure is increased by adding an inert gas?
- Volume is increased?
- A catalyst is added?

Lesson 9.4 – Equilibrium Calculations with ICE Tables

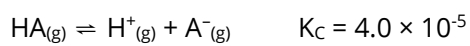
$$[A] = 2.0 \text{ M}, [B] = 0.5 \text{ M}$$

What are the concentrations of A, B, and C once the reaction reaches equilibrium?

Try it yourself!

$$[A] = 1.0 \text{ M}, [B] = 1.0 \text{ M}$$

What are the concentrations of A, B, and C once the reaction reaches equilibrium?



$$[HA] = 0.1 \text{ M}$$

What is the equilibrium concentration of H^+ ?

Lesson 9.5 – Calculating K_{sp} and Molar Solubility

1. Write out the balanced chemical equation
2. Write out the equilibrium expression, K_{sp}
3. Solve for whatever is missing

Molar Solubility

What is the molar solubility of $PbCl_2$? (K_{sp} is 2.5×10^{-4})

Calculating K_{sp}

When dissolved in water at a certain temperature, a certain ionic compound A_2B is found to have a solubility of 2×10^{-6} M. What is the K_{sp} of this compound?

Lesson 9.6 – Common Ion Effect and PrecipitationCommon Ion Effect

0.50M HCl is added to a solution of $PbCl_2$. What is the new molar solubility of $PbCl_2$? ($K_{sp} = 2.5 \times 10^{-4}$)

Precipitation

1×10^{-2} M $Cu(NO_3)_2$ was added to $NaIO_3$ to a final concentration of 6.0×10^{-3} . Does a precipitate form?
(K_{sp} of $Cu(IO_3)_2 = 3.6 \times 10^{-6}$)

Chapter 10: Acid-Base Equilibria and Titrations

Lesson 10.1 – Acid Definitions and Conjugate Acid/Base Pairs

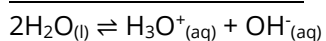
Acid-Base Definitions

	Arrhenius	Bronsted-Lowry	Lewis
Acid	H ⁺ donor in water	H ⁺ donor	Electron acceptor
Base	OH ⁻ donor in water	H ⁺ acceptor	Electron donors

Amphiprotic – describes a substance that can act as both a proton acceptor and donor (Arrhenius and Bronsted)

Amphoteric – describes a substance that can act as both an acid and a base (Lewis Acids too)

Autoionization of Water



$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1 \times 10^{-14} \quad \text{at } 25^\circ\text{C}$$

Conjugate Acid / Base Pairs

Conjugate Acid	Conjugate Base
	HSO_4^{-1}
	NO_3^{-1}
H_3O^+	
HSO_4^{-1}	

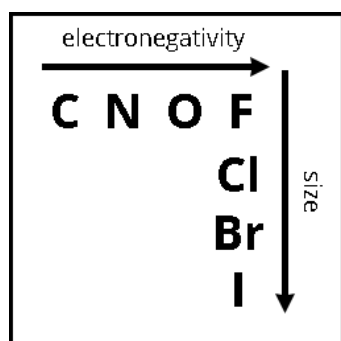
Lesson 10.2 – Strong Acids and Bases

- Strong acids/bases dissociate completely in water

Strong Acids	Strong Bases
HI	Group 1 metal hydroxides
HBr	$\text{Mg}(\text{OH})_2$
HCl	$\text{Ca}(\text{OH})_2$
HClO_3	$\text{Sr}(\text{OH})_2$
HClO_4	$\text{Ba}(\text{OH})_2$
H_2SO_4	
HNO_3	

Lesson 10.3 – Trends in Acid Strength

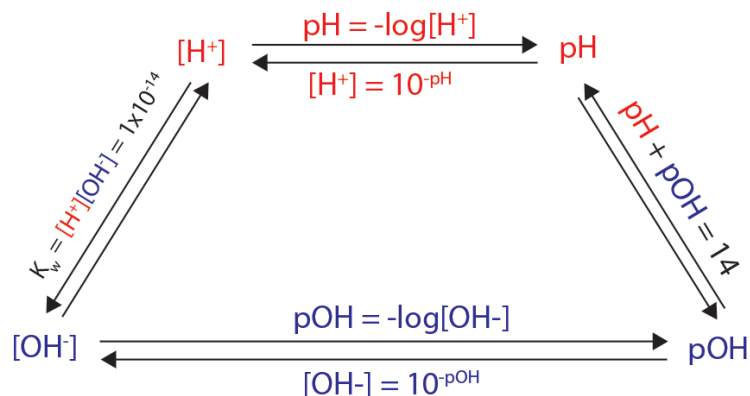
Binary Acid Trend (HCl, HBr, HI)

Oxoacid Trends (HClO₃, HBrO₄)

1. The more oxygens you have, the more acidic due to resonance (HClO₄ > HClO₃)
2. If the number of oxygens is the same, the more electronegative heteroatom, the more acidic (HClO₄ > HBrO₄)

Lesson 10.4 – pH and pOH Calculations

pH Scale	
$\text{pH} = -\log[\text{H}^+]$	$[\text{H}^+] = 10^{-\text{pH}}$
$\text{pOH} = -\log[\text{OH}^-]$	$[\text{OH}^-] = 10^{-\text{pOH}}$
$\text{pH} + \text{pOH} = 14$	$[\text{H}^+][\text{OH}^-] = 1 \times 10^{-14}$



[H ⁺]	pH	[OH ⁻]	pOH
	8		
			3
1x10 ⁻³ M			
		3.5x10 ⁻⁴ M	

Lesson 10.4 – pH and pOH Calculations (Continued)

Try it yourself!

What are the approximate pH and pOH values of an aqueous solution of 1.0×10^{-3} M HBr?

If a solution has a pH of 2.0, what are $[H^+]$ and $[OH^-]$?

Approximate the following pH values:

0.01 M HCl

0.01 M H_2SO_4

0.00320 M HNO_3

1×10^{-11} M HBr

0.01 M KOH

0.01 M $Ba(OH)_2$

Lesson 10.5 – Weak Acids and Bases

	Weak Acids	Weak Bases
Equation	$HA + H_2O \rightleftharpoons H_3O^+ + A^-$	$A^- + H_2O \rightleftharpoons HA + OH^-$
Equilibrium Expression	$K_a = \frac{[H_3O^+][A^-]}{[HA]}$	$K_b = \frac{[OH^-][HA]}{[A^-]}$
Shortcut	$[H^+] = \sqrt{K_a[HA]}$	$[OH^-] = \sqrt{K_b[A^-]}$

$$pK_A = -\log[K_A] \quad pK_B = -\log[K_b]$$

$$K_w = 1.0 \times 10^{-14} = [H^+][OH^-] = K_A \times K_B$$

$$pK_A + pK_b = 14$$

Larger K_a = smaller pK_a = stronger acid

Larger K_b = smaller pK_b = stronger base

Calculate the following:

What is the approximate pH of a 0.005 M solution of HF ($K_A = 5.0 \times 10^{-4}$)?

What is the approximate pH of a 0.02 M solution of NaF (the K_A of HF is 5.0×10^{-4})?

What is the approximate pH of a solution that is 0.02 M HF ($K_A = 5.0 \times 10^{-4}$) and 0.02 M NaF?

Lesson 10.6 – Neutralization Reactions and Normality

$$n_A M_A V_A = n_B M_B V_B$$

Calculate the following:

How many liters of 1.0 M HCl do you need to neutralize 2.0 L of 3.0 M NaOH?

How many liters of 1.0 M HCl do you need to neutralize 1.5 L of 3.0 M Ca(OH)₂?

How many liters of 1.0 **N** H₂SO₄ do you need to neutralize 1.5 liters of 3.0 N Sr(OH)₂?

How many liters of 1.0 M H₂SO₄ do you need to neutralize 1.5 liters of 3.0 M Sr(OH)₂?

Lesson 10.7 – Hydrolysis of Salts

Neutral cations	Neutral anions
Group 1 metals	Cl ⁻
Group 2 metals	Br ⁻
Metals with +1 charge	I ⁻
	NO ₃ ⁻
	ClO ₄ ⁻
	ClO ₃ ⁻

Label the following as acidic, basic, or neutral:



Lesson 10.7 – Hydrolysis of Salts (Continued)

Which of the following could you add to increase the solubility of CaF_2 ?

- A. $\text{Ca}(\text{OH})_2$
- B. NaNO_3
- C. KF
- D. HCl
- E. KNO_3

Which of the following could you add to increase the solubility of NH_4Cl ? (More than one may be correct)

- A. $\text{Ca}(\text{OH})_2$
- B. NaNO_3
- C. KF
- D. HCl
- E. KNO_3

Lesson 10.8 – Buffers and Henderson-Hasselbalch

Buffer – is a solution that resists pH change. It is made from a weak conjugate acid/conjugate base pair.

$$pH = pK_a + \log \frac{[A^-]}{[HA]}$$

Please identify the missing pK_a values. Then determine which of the listed acids could be used to prepare a buffer with a pH of 9.0.

Acid	K_a	pK_a
$\text{CH}_3\text{CH}_2\text{COOH}$	1.3×10^{-5}	
HCOOH	1.8×10^{-4}	
HNO_2	4.0×10^{-4}	
HCN	6.0×10^{-10}	

Which of the following substances could be combined to form a buffer solution?

- A. KI , HI
- B. KBr , HBr
- C. CuCl , HCl
- D. NaI , HI
- E. NaCH_3COO , CH_3COOH

Lesson 10.8 – Buffers and Henderson-Hasselbalch (Continued)

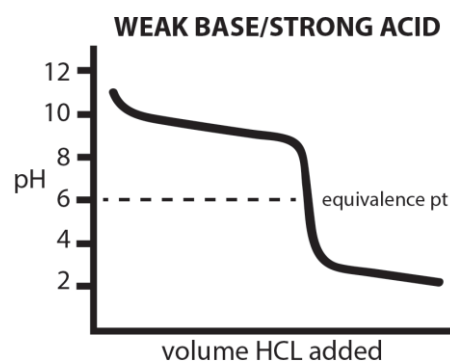
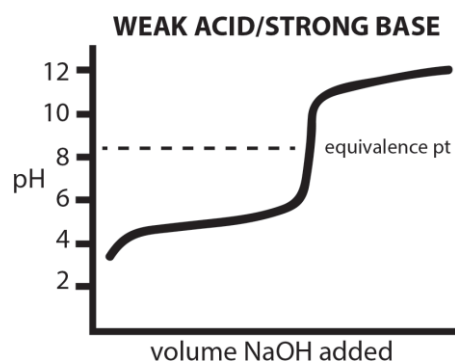
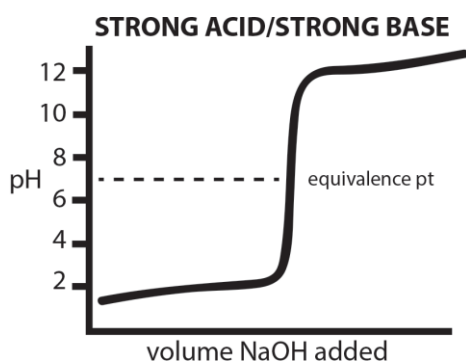
HF has a pK_A of 3.2. What is the pH of a solution comprised of 0.1 M HF and 0.1 M NaF?

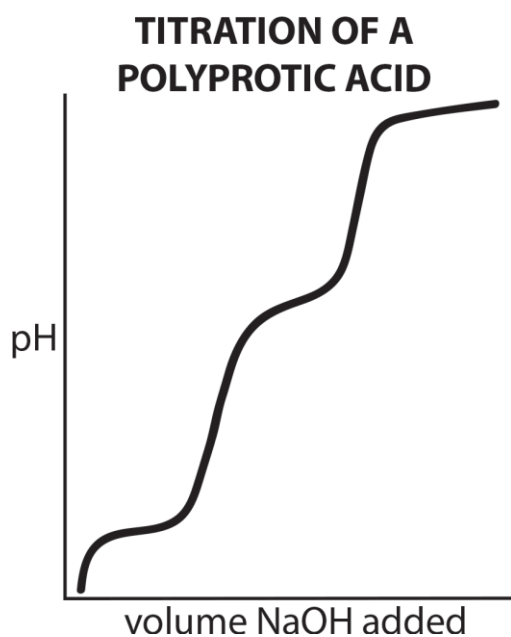
What is the pH of a solution comprised of 1.0 M HF ($pK_A = 3.2$) and 0.1 M NaF?

What is the pH of a solution comprised of 0.1 M HF ($pK_A = 3.2$) and 10 M NaF?

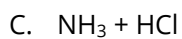
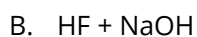
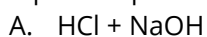
Lesson 10.9 – Titrations

Type of Titration	Example	Salt formed (Acidic/Basic)	pH at Equivalence pt.
Strong Acid/Strong Base	HCl + NaOH	NaCl (neutral)	7
Weak Acid/Strong Base	HF + NaOH	NaF (basic)	greater than 7
Weak Base/Strong Acid	NH ₃ + HCl	NH ₄ Cl (acidic)	less than 7



Lesson 10.9 – Titrations (Continued)

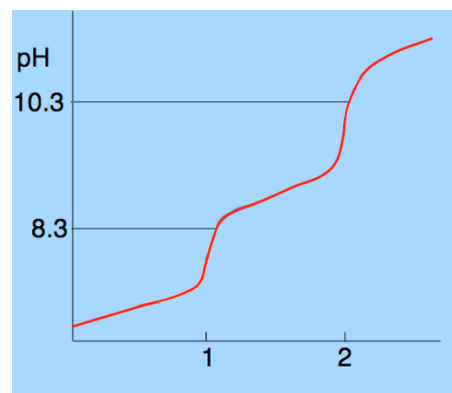
What will the pH at equivalence point be for each of the following titrations?



For a generic diprotic acid H_2A , whose titration curve is shown here:

- What are the values of $\text{pK}_{\text{A}1}$ and $\text{pK}_{\text{A}2}$?

- Which molecule predominates:
 - When $\text{pH} < \text{pK}_{\text{A}1}$?
 - When $\text{pK}_{\text{A}1} < \text{pH} < \text{pK}_{\text{A}2}$?
 - When $\text{pH} > \text{pK}_{\text{A}2}$?



Chapter 11: Thermodynamics and Thermochemistry

Lesson 11.1 – Laws of Thermodynamics and Definitions

Definitions

State Function – a function that only depends on the initial and final states of the system, not how it got there.
Almost everything in thermodynamics is a state function, *except* work (w) and heat (q).

System – The chemical reaction we are examining

Surroundings – Everything else around it.

Laws of Thermodynamics

1. (Conservation of Energy) - Energy cannot be created or destroyed. It can only be converted from one form to another.
 2. (Entropy increase) - The combined change in entropy of a system and its surroundings (the “universe”) must be positive
 3. (Perfect crystal at 0K = 0 entropy) - For a perfect crystal that’s been cooled to 0 K (–273 °C, called absolute zero), its entropy approaches zero.
-

Lesson 11.2 – Enthalpy

Enthalpy – the amount of heat-energy a substance contains

Exothermic – a process that gives off heat to its surroundings (negative ΔH). They feel hot to the touch.

Endothermic – a process that consumes heat from its surroundings (positive ΔH). They feel cold to the touch.

Solid → Liquid → Gas = Endothermic

Gas → Liquid → Solid = Exothermic

Lesson 11.3 – Enthalpies of Formation

Enthalpy of Formation – the ΔH of a chemical reaction forming one mole of a single substance from its parent elements in their standard states.

$$\Delta H_{\text{rxn}} = \sum nH_{\text{products}} - \sum nH_{\text{reactants}}$$

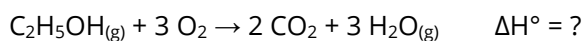
Standard States

The enthalpy of formation of any element in its standard state is 0.

Examples: Noble gases, H_2 , N_2 , O_2 , F_2 , Cl_2 , Br_2 , Hg , C (graphite), S_8

Example

Using the information in the table, calculate the ΔH_{rxn} for the following reaction:



ΔH_f° (kJ/mol)	
$\text{C}_2\text{H}_5\text{OH}(\text{g})$	-200
CO_2	-400
$\text{H}_2\text{O}(\text{g})$	-250

Lesson 11.4 – Thermodynamic Equations

$$\Delta E = q + w \quad w = -P\Delta V$$

+q (endothermic)	Heat is transferred to the system (from the surroundings)
-q (exothermic)	Heat is transferred to the surroundings (from the system)
+w	The surroundings do work on the system (compression)
-w	The system does work on the surroundings (expansion)

Lesson 11.5 – Heat Transfer and CalorimetryHeat Transfer

Conduction – heat transfer due to direct contact via molecular agitation within a material

Convection – heat transfer due to the motion of a fluid

Radiation – heat transfer via electromagnetic radiation

Calorimetry

Bomb calorimeter

$$q = -C_{\text{calorimeter}}\Delta T$$

$C_{\text{calorimeter}}$ = the calorimeter's specific heat

Example

You used a bomb calorimeter, whose heat capacity is 2000 cal/°C, to measure a banana's heat content (calories). During the banana's combustion, the temperature changed from 15 °C to 20 °C. What is this banana's q ?

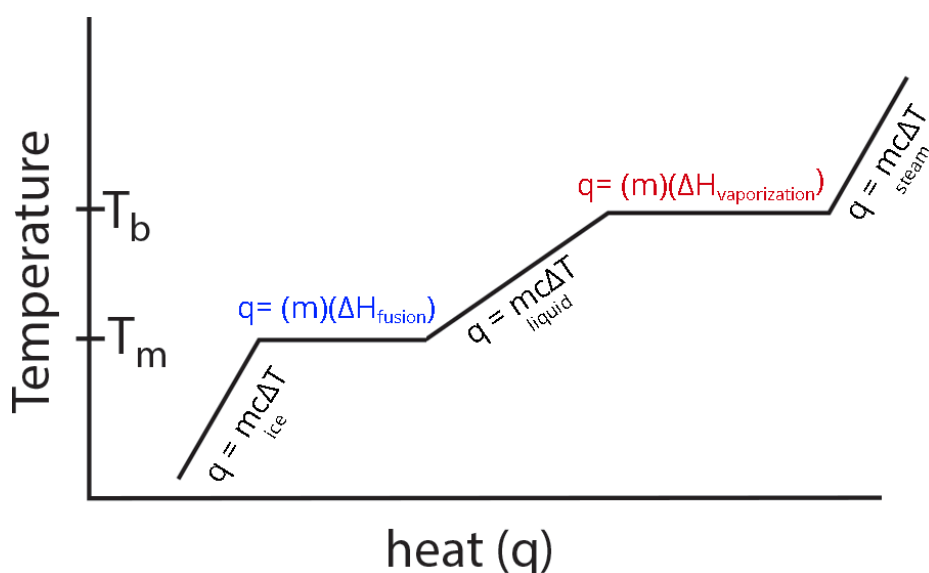
Lesson 11.6 – Heat Curves

$$q = mc\Delta T$$

c = specific heat

$$q = (m)(\Delta H_{\text{fusion}})$$

$$q = (m)(\Delta H_{\text{fusion}})$$



Lesson 11.6 – Heat Curves (Continued)

How much energy is required to increase the temperature of 36.0 g of H₂O from -20 °C to 50 °C, keeping in mind the following:

$$C_{\text{ice}} \approx 2.00 \text{ J/g}\times\text{C}$$

$$C_{\text{liquid water}} \approx 4.00 \text{ J/g}\times\text{C}$$

$$\Delta H_{\text{fusion}} \approx 6000 \text{ J/mol}$$

What is the final temperature (T_{Final}) of a 10.0 g ingot of silver that starts at 120 °C and is placed in 20.0 g of liquid H₂O at 40 °C, keeping in mind the following:

$$C_{\text{solid silver}} \approx 0.25 \text{ J/g}\times\text{C}$$

$$C_{\text{liquid water}} \approx 4.00 \text{ J/g}\times\text{C}$$

Lesson 11.7 – Entropy

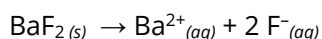
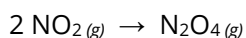
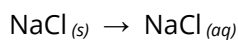
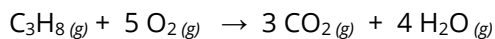
A measure of how disordered something is

$$S_{\text{solid}} < S_{\text{liquid}} < S_{\text{gas}}$$

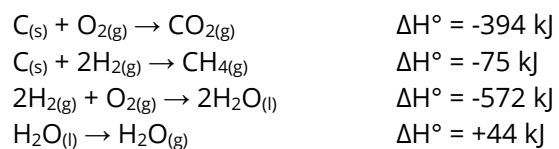
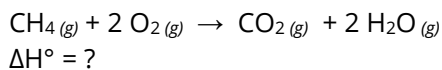
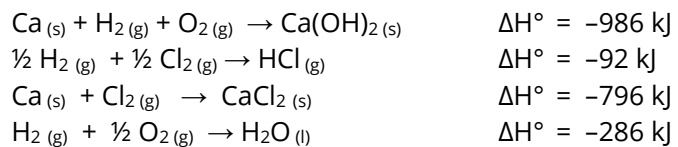
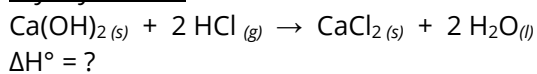
$$\Delta S = \sum nS_{\text{products}} - \sum nS_{\text{reactants}}$$

Example

Indicate if each of the following reactions or processes will have a positive or negative ΔS , and if they are entropically favorable or entropically unfavorable:



Predicting ΔS for Reactions	
$\Delta S_{\text{rxn}} > 0$	Sublimation, vaporization, fusion (going from solid to liquid to gas)
$\Delta S_{\text{rxn}} > 0$	Increase in number of moles of gas
$\Delta S_{\text{rxn}} > 0$	solid \rightarrow aqueous

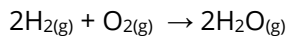
Lesson 11.8 – Hess’s LawTry it yourself!

Lesson 11.9 – Bond Dissociation Energy

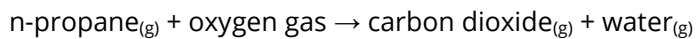
Making bonds = exothermic

Breaking bonds = endothermic

$$\Delta H = \Delta H_{\text{bonds broken}} - \Delta H_{\text{bonds formed}} = \Delta H_{\text{reactants}} - \Delta H_{\text{products}}$$

Given the following enthalpy data, what is the ΔH_{rxn} for the following reaction? $\Delta H = ?$

Bond	ΔH_{bond}
H-H	400 kJ
O=O	500 kJ
H-O	450 kJ

Try it yourself!Given the following enthalpy data, what is the ΔH_{rxn} for the following reaction? $\Delta H = ?$

Bond	ΔH_{bond}
C-H	410 kJ
C-C	350 kJ
O=O	500 kJ
C=O	800 kJ
H-O	450 kJ

Lesson 11.10 – Gibbs Free Energy

Gibbs free energy – a measure of spontaneity

ΔG	Meaning of ΔG Values Summary
negative	Spontaneous
positive	Nonspontaneous
zero	At equilibrium

$$\Delta G = \Delta H - T\Delta S$$

ΔH°	ΔS°	$-T\Delta S^\circ$	Reaction
-	+	-	Spontaneous at all temperatures
+	-	+	Nonspontaneous at all temperatures
-	-	+	Spontaneous at low temperatures
+	+	-	Spontaneous at high temperatures

Try it yourself!

1. Based on your experience with the following physical transformation (the melting of ice): $\text{H}_2\text{O}_{(s)} \rightarrow \text{H}_2\text{O}_{(l)}$

Is ΔH for this phase change positive or negative at room temperature (25 °C)?

Is ΔS for this phase change positive or negative?

Considering your answers to the previous questions, does this phase change occur spontaneously at room temperature? Is its ΔG positive or negative?

What could you do to make it nonspontaneous?

2. What can be said about a reaction that is known to be accompanied by an increase in entropy?

- The Gibbs free energy change of the reaction has a negative value at all temperatures.
- The reaction will be spontaneous at all temperatures.
- The reaction will be spontaneous above a certain temperature.
- The reaction will be spontaneous below a certain temperature.
- The reaction is spontaneous only at 25 °C.

Lesson 11.11 – Relating Gibbs Free Energy and K

ΔG°	K_{eq}	Significance
negative	$K > 1$	Reaction is spontaneous (products favored)
positive	$K < 1$	Reaction is nonspontaneous (reactants favored)
zero	$K = 1$	reactants and products equally favored

$$\Delta G = \Delta G^\circ + RT \ln Q$$

$$\Delta G^\circ = -RT \ln K_{eq}$$

Chapter 12: Electrochemistry and Oxidation-Reduction Reactions

Lesson 12.1 – Electrochemical Vocabulary

Oxidation – loss of electrons

Reduction – gain of electrons

LEO the lion says GER

OIL RIG

Oxidizing Agent (Oxidant) – what is reduced

Reducing Agent (Reductant) – what is oxidized

Anode – site of oxidation

Cathode – site of reduction

Lesson 12.2 – Oxidation States

<https://datbootcamp.com/blog/common-ions-you-need-to-know/>

1. Elements in their elemental form are zero oxidation state (Ex. Na, He, O₂, H₂, N₂, P₄, S₈, O₃ = 0 oxidation state).
2. Group 1 metals are +1, Group 2 metals are +2 in compounds.
3. Hydrogen is +1 when bonded to nonmetals and -1 when bonded to metals.
4. Transition elements' oxidation numbers must be determined from the other elements in the compound.
5. The most electronegative atoms get their typical oxidation state.
6. When assigning oxidation numbers, the last element assigned gets whatever number balances the charge for the entire compound.

Example

1. Assign oxidation states for all of the atoms to the following:

- Al₂O₃
- Na₂SO₃
- H₂O₂
- C₂H₃O₂¹⁻

2. For the following chemical reaction please assign the oxidation state for each element on both sides of the equation. Then answer the following questions: $\text{Ba}_{(s)} + \text{Cl}_{2(g)} \rightarrow \text{BaCl}_{2(s)}$

Which reactant gets oxidized?

What is the reducing agent?

Which reactant gets reduced?

What is the oxidizing agent?

Lesson 12.3 – Balancing Redox ReactionsBalancing under Acidic Conditions

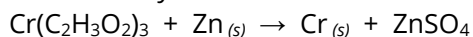
1. Identify what is getting oxidized and what is getting reduced.
2. Separate the oxidation reaction from the reduction reaction. (These two reactions you've now written are called *half-reactions*.)
3. Balance all the atoms that are NOT oxygen or hydrogen.
4. Balance oxygen atoms by adding H₂O to whichever side is needed.
5. Balance hydrogen atoms by adding H⁺ to whichever side is needed.
6. Add e⁻ wherever necessary, to balance the charge on each side of the equation.
7. As needed, add integers to your *half-reactions* to make the number of moles of e⁻ equal in each of your *half-reactions*. This final, balanced number of e⁻ is the number of overall electrons-per-mole that get transferred in this particular *redox reaction*.

Balancing under Basic Conditions

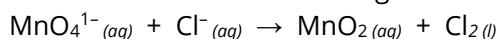
1. First, follow the same steps as for acidic conditions.
2. For each H⁺ you added to your *half-reactions*, now add the same number of OH⁻'s to both sides of the equation. Now, wherever an H⁺ and an OH⁻ appear on the same side an equation, combine them to form H₂O's. Eliminate excess H₂O.
3. Add e⁻ wherever necessary, to balance the charge on each side of the equation.
4. As needed, add integers to your *half-reactions* to make the number of moles of e⁻ equal in each of your *half-reactions*. This final, balanced number of e⁻ is the number of overall electrons-per-mole that get transferred in this particular *redox reaction*.

Examples

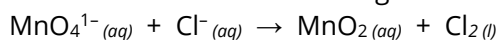
1. How many electrons are transferred in this reaction?



2. Please balance the following redox reaction under **acidic** conditions and find the number of e⁻ transferred:



3. Please balance the following redox reaction under **basic** conditions and find the number of e⁻ transferred:



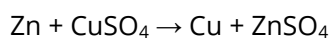
Lesson 12.4 – Voltaic Cells

1. Anode = always site of oxidation
2. Cathode = always site of reduction
3. Electrons always flow from anode to cathode
4. In voltaic/galvanic cells, the anode is negative and the cathode is positive
5. In electrolytic cells, the anode is positive and the cathode is negative
6. The salt bridge lets counter anions flow into the anode, and cations to the cathode
7. The anode metal loses mass, while the cathode metal gains mass

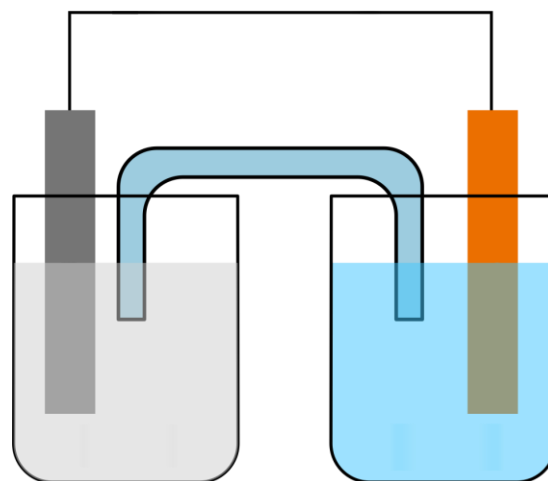
Voltaic (Galvanic) Cells
Spontaneous ($-\Delta G$, $+E_{\text{cell}}$)
Produce electricity (batteries)
Anode(-) \rightarrow cathode(+)

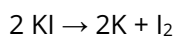
Example

Using the diagram to the right for the following reaction in a galvanic cell, please indicate or answer each of the following:



- Identify the elements being oxidized and reduced.
- What flows through the wire?
- What flows through the salt bridge?
- Label the flow of electrons in the correct direction.
- What is each metal made of at the anode and cathode? What is each solution made of?
- What is the purpose of the counterion?
- What is a salt bridge?



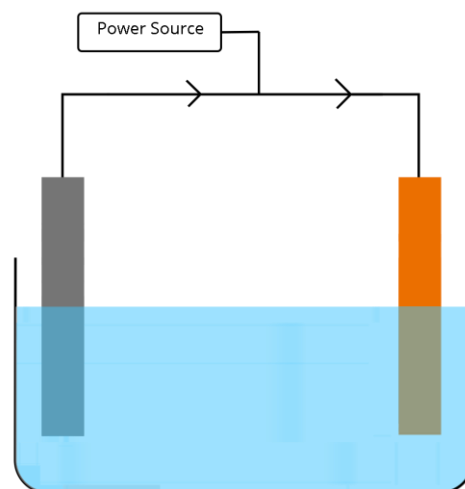
Lesson 12.5 – Electrolytic Cells and ElectrolysisMolten Electrolysis

- Under molten electrolysis, there is no salt bridge.
- There is a sea of molten K^+ and I^- ions.
- Electrons ALWAYS flow from anode to cathode.
- Electrolytic cells need a power source.
- Cations flow towards the cathode to get reduced.
- Anions flow towards the anode to get oxidized.

Aqueous Electrolysis

- Under aqueous conditions you have two competing reactions (reaction of KI and reaction of H_2O)
- You need to use a table of reduction potentials to find which one will happen more easily (next lesson).

Electrolytic Cells
Nonspontaneous ($+\Delta G$, $-E_{\text{cell}}$)
Consume electricity (makes elements)
Anode(+) \rightarrow cathode(-)



Lesson 12.6 – Reduction Potentials

$$E_{\text{cell}} = E_{\text{reduction}} + E_{\text{oxidation}} \quad (\text{or}) \quad E_{\text{cell}} = E_{\text{cathode}} + E_{\text{anode}}$$

What is the strongest oxidizing agent?

What is the strongest reducing agent?

What are some pairs that react spontaneously?

Reduction Potentials

$K^+ + 1e^- \rightarrow K$	-2.93V
$Na^+ + 1e^- \rightarrow Na$	-2.71V
$Al^{3+} + 3e^- \rightarrow Al$	-1.66V
$Mn^{2+} + 2e^- \rightarrow Mn$	-1.18V
$Zn^{2+} + 2e^- \rightarrow Zn$	-0.76V
$Cr^{3+} + 3e^- \rightarrow Cr$	-0.74V
$Fe^{2+} + 2e^- \rightarrow Fe$	-0.44V
$Ni^{2+} + 2e^- \rightarrow Ni$	-0.25V
$2H^+ + 2e^- \rightarrow H_2$	0.00V
$Cu^{2+} + 2e^- \rightarrow Cu$	+0.34V
$I_2 + 2e^- \rightarrow 2I^{1-}$	+0.54V
$Ag^+ + 1e^- \rightarrow Ag$	+0.80V

When Mn and Zn metals react in a galvanic cell according to the equation below, will the reaction be spontaneous or nonspontaneous? What will be the products of this reaction? $Mn + Zn \rightarrow ?$

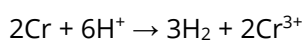
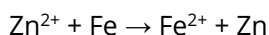
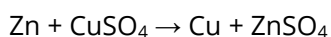
When Na metal and Fe^{2+} cation react in a galvanic cell according to the equation below, will the reaction be spontaneous or nonspontaneous? What will be the products of this reaction? $Na + Fe^{2+} \rightarrow ?$

What are the products of the electrolysis of $KI_{(aq)}$?

$K^+ + 1e^- \rightarrow K$	-2.93V
$I_2 + 2e^- \rightarrow 2I^-$	+0.54V
$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$	-0.80V
$H_2O \rightarrow O_2 + 4H^+ + 4e^-$	-1.28V

Lesson 12.7 – Standard Reduction Potentials

What is the standard cell potential, or E° , for the following reactions?

**Reduction Potentials**

$\text{K}^+ + 1\text{e}^- \rightarrow \text{K}$	-2.93V
$\text{Na}^+ + 1\text{e}^- \rightarrow \text{Na}$	-2.71V
$\text{Al}^{3+} + 3\text{e}^- \rightarrow \text{Al}$	-1.66V
$\text{Mn}^{2+} + 2\text{e}^- \rightarrow \text{Mn}$	-1.18V
$\text{Zn}^{2+} + 2\text{e}^- \rightarrow \text{Zn}$	-0.76V
$\text{Cr}^{3+} + 3\text{e}^- \rightarrow \text{Cr}$	-0.74V
$\text{Fe}^{2+} + 2\text{e}^- \rightarrow \text{Fe}$	-0.44V
$\text{Ni}^{2+} + 2\text{e}^- \rightarrow \text{Ni}$	-0.25V
$2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$	0.00V
$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$	+0.34V
$\text{I}_2 + 2\text{e}^- \rightarrow 2\text{I}^-$	+0.54V
$\text{Ag}^+ + 1\text{e}^- \rightarrow \text{Ag}$	+0.80V

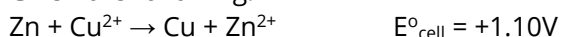
Lesson 12.8 – Non-standard Cell Potentials and Nernst EquationNernst Equation

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0592}{n} \log(Q)$$

1. Shift to the right = increase E_{cell}
2. Shift to the left = decrease E_{cell}

Example

Given the following:



In which direction will the reaction shift if $[\text{Cu}^{2+}]$ is increased to 10 M?

How will this change in $[\text{Cu}^{2+}]$ affect E_{cell} ?

Use the Nernst equation to solve for E_{cell} under these new conditions.

Lesson 12.9 – Quantitative Electrochemical Calculations

$$\frac{(I)(t_s)}{(n)(F)} = \text{moles of product}$$

I = current (amps)

t_s = time (seconds)

n = number of electrons transferred

F = Faraday's Constant (96,500 C / mol e^-)

If a 20-amp current passes through molten BaCl_2 for 30 minutes, how many grams of Ba are produced?

What amount of time (in seconds) would it take to plate out 0.50 kg of Ag from molten AgCl with a current of 100 amps? (atomic weight of Ag = 100 g/mol)

Chapter 13: Nuclear Reactions

Lesson 13.1 – The Nucleus and Nuclear Particles

The Nucleus

Nucleus – made of protons and neutrons

Atomic number – number of protons

Atomic weight vs. Mass number – atomic weight is an average of all isotopes of an atom

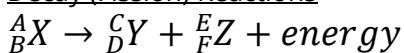
Particle	Symbol
Alpha Particle	${}^4_2\text{He}$ or ${}^4_2\alpha$
Neutron	${}_0^1\text{n}$
Proton	${}_1^1\text{H}$ or ${}_1^1\text{p}$
Beta Particle (high-speed electron)	${}_{-1}^0\text{e}$ or ${}_{-1}^0\beta$
Positron	${}_{+1}^0\text{e}$
Gamma particle	${}^0_0\gamma$

↑ increasing mass ↓ increasing penetration

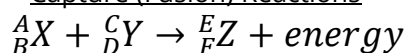
Lesson 13.2 – Nuclear Reactions

Mass is always lost in nuclear reactions.

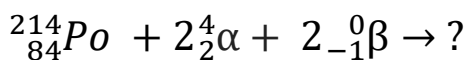
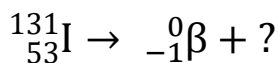
Decay (Fission) Reactions



Capture (Fusion) Reactions

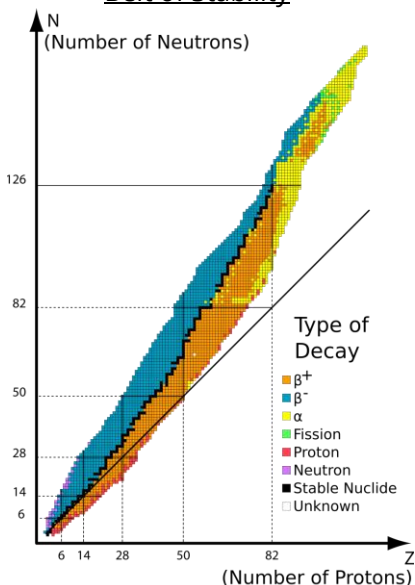


Examples



Lesson 13.3 – Nuclear Stability and Radioactive Decay

Belt of Stability



Radioactive Decay

Route of Decay	Nuclear particle	Result	Likely For?
α -decay	${}^4_2\alpha$ product	Reduces mass #	Large nuclei
β decay (β emission)	${}_{-1}^0\beta$ product	neutron \rightarrow proton	N/Z ratio too high (too many neutrons)
β^+ decay (positron emission)	${}_{+1}^0\beta$ product	proton \rightarrow neutron	N/Z ratio is too low (too many protons)
Electron capture	${}_{-1}^0\beta$ reactant	proton \rightarrow neutron	N/Z ratio is too low (too many protons)
γ decay	${}^0_0\gamma$ product	no change	unpredictable

Lesson 13.3 – Nuclear Stability and Radioactive Decay (Continued)Example

Oxygen-17 and a proton combine and then undergo radioactive decay to emit a nitrogen-14 atom and another particle. What type of radioactive decay is oxygen-17 undergoing in this scenario?

Lesson 13.4 – Kinetics of Radioactive Decay

Always 1st order kinetics. Half-life is constant.

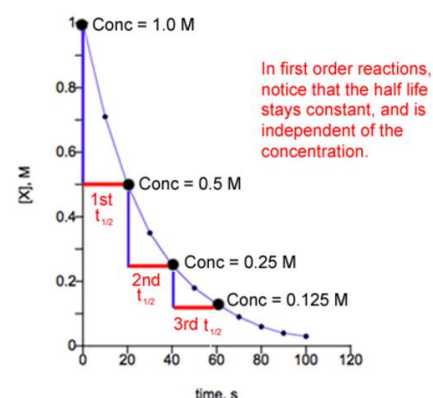
$$N = N_0 e^{-kt} \quad \text{or} \quad \ln(N) = \ln(N_0) - kt$$

$$t_{1/2} = (0.693/k)$$

Examples

Given a half-life of 30 minutes and a starting sample of 100 grams, how many grams would you have left after 2 hours?

$$\text{half-life of first order reaction} = \frac{\ln 2}{k} = \frac{0.693}{k}$$



Given a half-life of 15 minutes, how long would it take for 87.5 grams of your 100-gram sample to decay?

Lesson 13.5 – Fission, Fusion, Nuclear Binding EnergyFission and Fusion

Fission – larger nuclei decay into smaller nuclei (atomic bombs, nuclear reactors)

Fusion – smaller nuclei combine to form larger nuclei (stars)

Nuclear Binding Energy

Mass defect – a nucleus always weighs a little less than its constituent nucleons

$$E = \Delta mc^2 \quad \Delta m \text{ in kg}$$

⁵⁶Fe has highest nuclear binding energy per nucleon (most stable nucleus)

Example (Use a calculator)

A proton's mass is 1.00728 amu. A neutron's mass is 1.00867 amu. What is the mass defect of a ¹⁷⁸Au nucleus, if a ¹⁷⁸Au nucleus weighs 177.9760 amu?

What is the binding energy (in J) of a ¹⁷⁸Au nucleus?