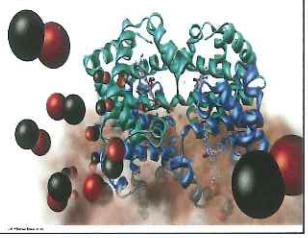
Chapter 1

"Matter, Measurement, and Problem Solving"



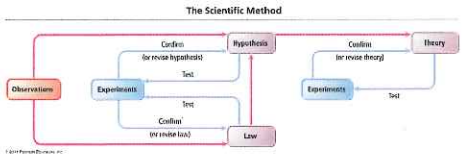
Matter

- All matter is composed of atoms
- Atoms: the smallest identifiable piece of matter
- Molecules: two or more atoms joined by a covalent bond

The Scientific Approach

- Philosophers: use reasoning and thinking to understand the universe
- Scientists: use empirical knowledge gained through observation and experimentation to understand the universe

The Scientific Approach

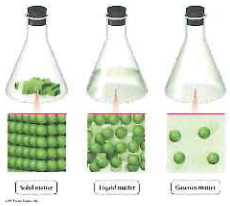


Observation → of Law
↳ Hypothesis ↳ test
↳ Test ↳ Confirm
↳ Confirm
↳ Theory

Classification by State

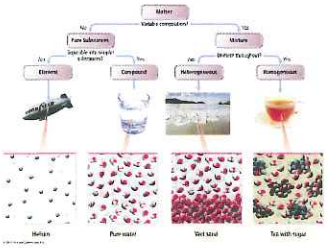
fixed Volume by Conservation of Volume

- Solid
- Liquid
- Gas



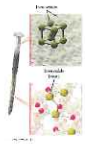
Solid – fixed Volume & Shape
Liquid – fixed Volume & Variable Shape
Gas – Variable Shape & Volume

Classification by Composition



matter → Mixture
↳ Substances ↳ Hetero
↳ Elements Homog
Compounds

Changes and Properties

	Physical	Chemical
Changes	Alters the state or appearance	Alters the substance
Examples	Water boiling	Iron rusting 

No new substance formed
↳ Physical Change
New Substance formed
↳ Chemical Change

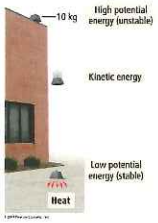
Changes and Properties

	Physical	Chemical
Properties	Observable without changing the substance	Observable by changing the substance
Examples	Density	Oxidation

Always do Physical test before Chemical tests

Energy: the capacity to do work

- Kinetic energy:
  - Energy of motion
  - Heat, sound, light, gas expansion, movement of electrons.
- Potential energy:
  - Energy of position
  - Stored in compounds.



Kinetic ⟷ Potential

Heat - Kinetic Energy
Sound - Chemical Kinetic

Potential - Stored energy in a compound. Like gasoline

we are given conversion factors!

SI Units of Measurement

Length	Meter	m
Mass	Kilogram	kg
Time	Second	s
Temperature	Kelvin	K

Mass - we use grams
Celsius - measurement in lab

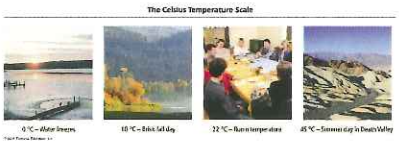
Temperature

°F ⟷ °C ⟷ K

°C = (°F - 32) / 1.8 | K = °C + 273.15

Dont count the 1.8 for sig figs since its part of equation

Temperature



Ex; °F → °C
52.3°F → °C
(52.3 - 32) ÷ 1.8
= 11.2777 = 11.3°C

Example: Temperature

$-34.4\,°C \rightarrow °F$

$(-34.4 * 1.8) + 32$

$= -29.9\,°F$

Derived Units

- Volume:
  - L x w x h
  - 1cm x 1cm x 1cm
  - $1cm^3 = 1mL$
- The Liter is the base unit of volume = L

$1cm^3 = 1mL$

Volume is considered a derived unit

Derived Units

- Density:
  - Defined as mass per unit volume
  - Solid: g/cm^3
  - Liquid: g/mL
  - Gas: g/L

Gas has to be in g/L always

Example: Density

TABLE 1.4 The Density of Some Common Substances at 20 °C

Substance	Density (g/cm^3)
Charcoal (from oak)	0.57
Ethanol	0.789
Ice	0.917 (at 0 °C)
Water	1.00 (at 4 °C)
Sugar (sucrose)	1.58
Table salt (sodium chloride)	2.16
Glass	2.6
Aluminum	2.70
Titanium	4.51
Iron	7.86
Copper	8.96
Lead	11.4
Mercury	13.55
Gold	19.3
Platinum	21.4

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This chart of conversions will be given on exam

Example: Density calculations

mass = 0.456 lb

Volume = 6.32 cm → $(6.32 cm)^3$

Volume = 252.436

mass = .456 lb * $\frac{453.6 g}{1 lb}$ = 206.8416

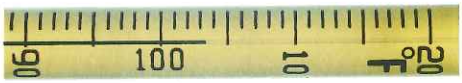
$d = \frac{206.8416}{252.436} = \boxed{0.819 \frac{g}{cm^3}}$

$d = \frac{m}{V}$ 453.6 g = 1 Lb

Dont truncate until the end

Significant Figures

- Based on the digit of uncertainty
- Mentally you divide the area between two marks by ten
  - 12.3 means the reading was between the 12 and the 13 and closer to the 12
  - 12.29 means the reading was between the 12.2 and 12.3 and very close to the 12.3



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You always need the digit of uncertainty in every measurement

→ Guess is 103.4

Guessed digit is the .4

Significant Figures: Rules

- All non-zeros are significant
  - 1987.67
- Interiors zero are significant
  - 980,078.008
- Leading zeros are not significant
  - 0.0000349
- Trailing zeros may be significant
  - 19800 (3) 1700. (4) 12.000 (5)

NO Yes Yes

Significant Figures: Rules

- Exact numbers have unlimited significant figures
- Examples:
  - Counting numbers
  - Definitions → Ex: 60sec in 1min
  - Integer values in equations

Conversions have Sig Figs between other systems & so do measurements

Significant Figures: Calculations

- Multiplication and Division: the least number of sig. figs. in the calculation determine the number of digits in the answer.
- Addition and Subtraction: the measurement with the least precision determines the precision of the answer. →

Mult-Divide: get the smallest sig figs in Calculation & Truncate

Least Precise = number with the least number of decimals

Ex:

$$\begin{array}{r} 12{,}300 \text{ ft} \\ + \quad 44 \text{ ft} \\ \hline 12{,}344 \text{ ft} \end{array}$$

↓

$\boxed{12{,}300 \text{ ft}}$

Significant Figures: Calculations

- Multiplication and Division: Example
  - 5.02ft x 89.665ft x 0.10ft =
- Addition and Subtraction: Example
  - 5.74g + 0.823g + 2.651g =

At each step please truncate before performing the next operation

Ex: (1.428 − 1.08) ÷ 0.288
= 0.348 ÷ 0.288
= 0.35 ÷ 0.288 = 1.2

Do not double round!
Ex: 2.349 → 2.3 not 2.4

Dimensional Analysis

- Used in problem solving
- Arrange conversion factors such that given units cancel each other
- 9ft x 1yd/3ft = 3yd

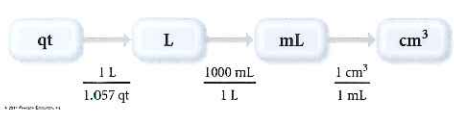
1 yard = 3 feet so

1 yd/3 ft or 3 ft/1 yd

1 inch = 2.54 centimeters
1 in/2.54 cm or 2.54 cm/1 in

Conceptual plan

- A conceptual plan is a visual outline that show the strategic route required to solve a problem



if you have no idea how to proceed on exam at least do a concept map

Concept map can get you some points

The End

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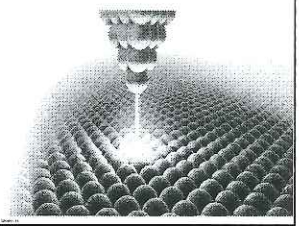
Chapter 2

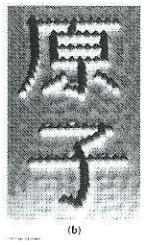
"Atoms and Elements"



if mentioned in the slide you need to know who they are & what they did

Early philosophy of Matter

- Democritus and Leucippus believed:
  - matter had an ultimate, tiny, indivisible particle
- Plato and Aristotle believed:
  - matter was infinitely divisible



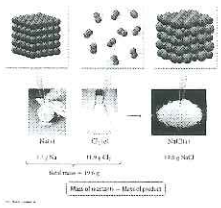
Dem: Says if you cut matter enough you'll end up with an uncuttable piece called the atom

Revolution

- In the late 16th century, the scientific approach to understanding was established
- Over the next 150+ years Aristotle view of matter could not easily explain the observation being made

Revolution: New Laws

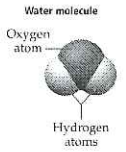
- Law of Conservation of Mass:
  - Lavoisier
  - In a chemical reaction matter is neither created nor destroyed



This is the start of experimental change

Revolution: New Laws

- Law of Definite Proportions:
  - Proust
  - All samples of the same compound have the same proportions of their constituent elements



if you look at the ratio of compounds you will notice that every molecule (like H_2O) has the same ratio throughout

Revolution: New Laws

- Law of Multiple Proportions:
  - Dalton
  - When two elements combine to form two different compounds the masses that combine can be expressed in small whole number ratios

2 Oxygen
1 Carbon = Carbon Dioxide

1 carbon
1 Oxygen = carbon monoxide
↓
deadly

two elements at different ratios mixed form something new

Dalton's Atomic Theory

Protons
Neutr. —

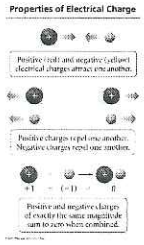
- 1. Each element is composed of tiny, nope indestructible particles called atoms

isotopes —

- 2. All atoms of a given element have the same mass and other properties that distinguish them from other elements
- 3. Atom combine in small whole number ratios to form molecules of compounds
- 4. In a chemical reaction, atoms of one element cannot change into an atom of another element

elements will not change in reaction!

Notes on Charges

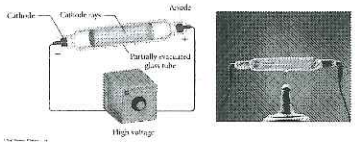


Pos repel Pos
Neg repel Neg

Pos & Neg particle of same magnitud cancel each other

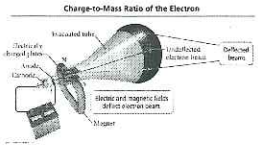
Thompson
The Discovery of Electrons

- A cathode ray tube is a glass tube containing metal rods from which almost all the air has been evacuated



The Discovery of Electrons

- J.J. Thomson believed that the cathode ray was composed of tiny particles with an electrical charge



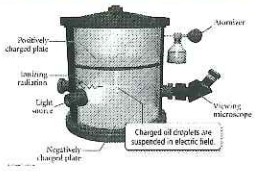
Applied magnetic field & show you can bend it

The Discovery of Electrons

- Thomson's conclusions:
  - The cathode rays are made of tiny negatively charge particles
  - Every material tested contained these same particles
  - Thomson believed that these particles were the ultimate building blocks of matter
  - These particle became known as electrons

The Discovery of Electrons

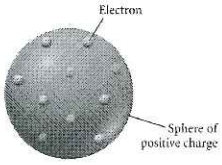
- Millikan's Oil Drop Experiment:
  - Determined the charge of an electron



Showed the charge of the electron

The Structure of the Atom

- Thomson:
  - The structure of the atom is like plum pudding
  - Negatively charged particle is a sphere of positive charge



Plum-pudding model

The Structure of the Atom

- The Plum Pudding model:
  - The mass of the atom is due to the mass of the electrons within
  - The atom is mostly empty space because negative particles repel

Problem: You need a whole lot of $\bar{e}$ to make up mass of bigger elements

Radioactivity

- Curie and Becquerel discovered that certain elements would constantly emit small, energetic particles and rays
- These energetic particles could penetrate matter

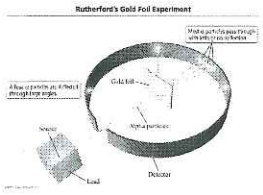
Radioactivity

- Rutherford discovered there were three type of radioactive emissions
  - Alpha:
    - mass of 4x a hydrogen atom and + charge
  - Beta:
    - mass of 1/2000th of a H atom and - charge
  - Gamma
    - Energy rays not particles

Disproves Thompsons model

Rutherford's Experiment

- **Hypothesis:** shoot alpha particles at a very thin sheet of matter and show that they all pass through

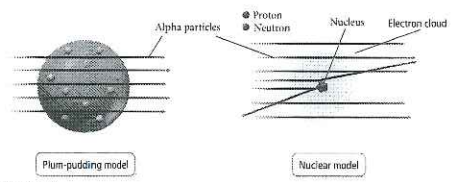


he expected for alpha particles to go through the paper thin gold but some reflected/bounced back!

Rutherford's Experiment

- Results:
  - 98% of the particles went straight through
  - About 2% of the particles went through but were deflected by large angles
  - About 0.01% of the particles bounced off the gold foil

Rutherford's Experiment



Positive rays hits the protons & reflects back or veer off

Rutherford's Experiment

- Conclusions:
  - The atom contains a tiny dense center called a nucleus
  - The atom is mostly empty space
  - The electron are dispersed in the empty space around the nucleus
  - The nucleus is positively charged
  - The nucleus of the atom must have a particle of the same amount of charge but opposite in sign

Lec-04

Fundamental Subatomic Particles

TABLE 2.1 Subatomic Particles

	Mass (kg)	Mass (amu)	Charge (relative)	Charge (C)
Proton	1.67262×10^{-27}	1.00727	+1	$+1.60218 \times 10^{-19}$
Neutron	1.67493×10^{-27}	1.00866	0	0
Electron	0.00091×10^{-27}	0.00055	−1	-1.60218×10^{-19}

Elements

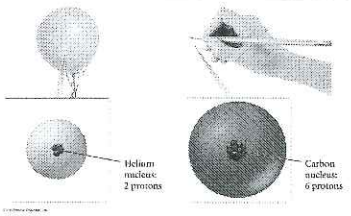
- Each element has a unique number of protons in its nucleus
- The number of protons is called the atomic number
- Each element has a unique name and symbol

17
Cl
35.45
chlorine

Number Protons also called the Z number

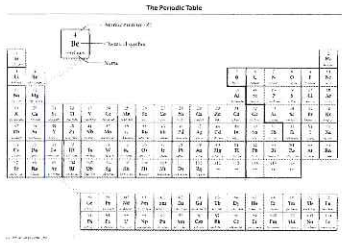
Elements

The Number of Protons Defines the Element



The Periodic Table of the Elements

The Periodic Table



Protons, Electrons and Neutrons

- Protons: determines the element
  - Number of protons = atomic number
- Electrons: determines the reactivity
  - If neutral # of electrons = # of protons
  - If charged: the sum of the charged particle must equal the charge
- Neutrons: determines the nuclear stability
  - The atomic mass (whole number) - the # of protons = the # of neutrons

^{14}C , $^{14}_{6}C$ C-14

The closer neutrons & Protons have a 1 to 1 ratio, the less radioactive it is

Protons, Electrons and Neutrons

- Ions: different number of electrons than neutral
  - Cations: have a positive charge, less electrons than neutral
  - Anions: have a negative charge, more electrons than neutral
- Isotopes: different number of neutrons
  - Can be stable or radioactive
  - Have different atomic mass
  - The elemental mass on the periodic table is usually close to the most common isotope
  - The mass on the periodic chart is calculated from the sum of all the weighted masses of the naturally occurring isotopes

Cation = Positive
The letter t in the word looks like a + sign

Examples: P, E, and N

Protons in S^{2-}?
Electrons in S^{2-}?

Ga^{+3} = P: 31, E: 28, N: 70-31 = 39

Protons = 16
Electrons = 18

Boron Average weight = 10.81

B-10 weight = 10.013
B-11 weight = 11.009

what is the percent composition for each?

formula = (Percent$_1$ * Weight$_1$) + (Percent$_2$ * weight$_2$) = Average weight

$$= (P_1 * 10.013) + (P_2 * 11.009) = 10.81$$

$(100 - P_2)$

$$= ((100 - P_2) * 10.013) + (P_2 * 11.009) = 10.81$$

Exam Example

Example: calculation of atomic mass

- If copper is 69.17% Cu-63 with a mass of 62.9396amu and the rest Cu-65 with a mass of 64.9278amu, what is copper's atomic mass?

Cu-65 = 100 - 69.17 = 30.83 %

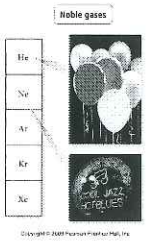
Cu-63 0.6917 * 62.9396 = 43.54
Cu-65 0.3083 * 64.9278 = + 20.02
63.56

Metallic Grouping

- Metals: shiny, ductile, malleable, conducts heat and electricity well
- Nonmetals: dull, brittle, poor conductor of heat and electricity
- Metalloids: semimetals, semiconductors, show properties of both metals and nonmetals

Some Important Families

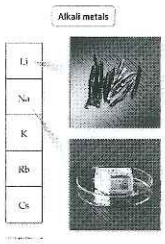
- Noble gases:
  - Non-reactive



All non-reactive except Rn

Some Important Families

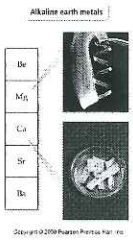
- Alkali metals:
  - Very reactive metals



All very reactive except Fr & H

Some Important Families

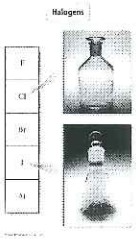
- Alkaline Earth metals:
  - Reactive metals



All reactive (although less reactive) except Ra

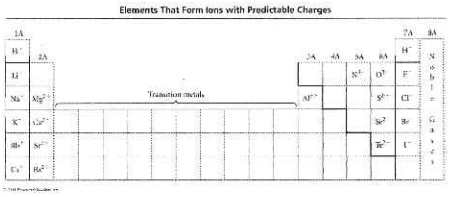
Some Important Families

- Halogens:
  - Reactive nonmetals



They are all diatomic
Typically forms acids

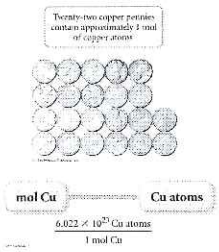
Patterns and Ions



H^- = Hydride ion
Be = ignore, it likes covalent bonds

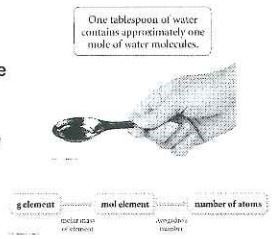
Counting Atoms by the Mole

- A mole is defined as a unit having 6.022×10^{23} particles
- A mole can be used to determine the number of atoms or molecules in a sample.



Counting Atoms by the Mole

- Recall: To find the number of moles in a sample divide the number of grams by the molar mass
- To find the number of particles in that sample multiply the number of moles by the Avogadro's number



Example: Counting Atom

- Determine the number of copper atoms in a penny weighing 10.9g.

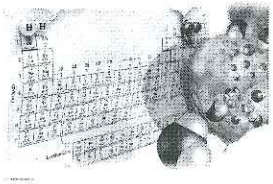
$\frac{\text{grams}}{\text{molar mass}} = \text{mols} * \frac{6.022*10^{23}\text{ atoms}}{\text{mol}}$

Avogadros # = $6.022*10^{23}\ \frac{\text{atoms}}{\text{mol}}$

The End

Chapter 3

"Molecules, Compounds, and Chemical Equations"



Compounds

- Compound: The chemical combination of two or more atoms of different elements
- The properties of the compound are totally different from the combining elements

Bonds in Compounds

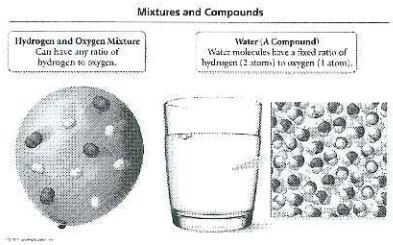
- Covalent Bonds: results when two atoms share electrons
- Generally between two nonmetals

- Ionic Bonds: results from the transfer of electrons or two ions coming together
- Electrostatic interaction holds the ions together together
- Generally between a metal and a non-metal

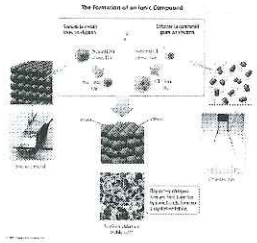
* Polyatomic Ions are bonded covalently but exhibit a charge

Ex: Na_2SO_4
both ionic & covalent bonds

Formation of Water



Formation of Sodium Chloride



Types of Formulas

- Empirical Formula:
- Describes the elements in the compound
- Tells the smallest whole number ratio of the elements to each other

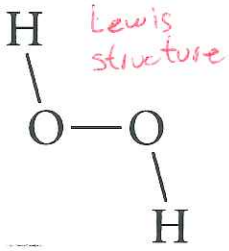
- Molecular Formula:
- Describes the elements in the molecule
- Tells the number of atoms in the molecule

↳ no ionics

Ex: H_2O_2 molecular
HO Empirical

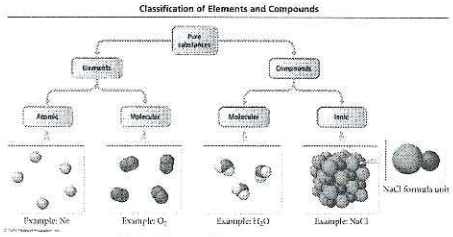
Types of Formulas

- Structural Formula:
- Describes the types of elements in the molecule
- Describe the arrangement of the atoms
- Shows the bonds between the atoms



Lewis structure

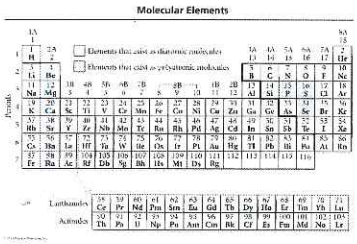
Classifying Materials



↳ Diatomic

BrINClHOF

Molecular Elements



Naming

- Review the naming of:
  - Molecules
  - Ionic compounds
  - Acids

Molecules – Uses Prefixes & use "ide" ending if not metallic
Ionic – Type I – group I & II A
Type II – need roman numerals
Acids – Non Oxy – uses "hydro" prefix
Oxy Acids – Oxygen present & Polyatomic
ate – ic
ite – ous

Formula Mass

- The mass of an single molecule or a formula unit
- AKA: molecular mass
- Calculation of: the sum of all the atomic masses

Molar Mass

- A mole contains 6.022 X10^{23} particles → Avogadros
- 1 mole of any compound has mass equal to the sum of the atomic masses with units of grams
- The unit of molar mass is g/mol

Example using Molar Mass

- Find the number of molecules of CO_2 in 11.4g of dry ice.

$C = 12.01 + O_2 = 16.00 \times 2$

$CO_2 = 44.01$ g/mol

$11.4g * \frac{mol}{44.01 g} * \frac{6.022 * 10^{23} molec}{mol}$

$= 1.56 * 10^{23}$ CO_2 molecules

Atoms? multiply response by subscript of the element you need

Percent Composition

- Percent of part per hundred
  - (part / whole) x 100 =
- It is a mass percent
- The percent composition can be determined from
  - The formula of the compound
  - The experimental mass analysis of the compound

Example Percent Composition

- Determine the mass of percent of each element in $C_2Cl_4F_2$ = 203.8

C $\frac{12.01}{203.8} * 100\% = 5.893 * 2 = 11.79\%$

Cl $\frac{35.45}{203.8} * 100\% = 17.39 * 4 = 69.58\%$

F $\frac{19.00}{203.8} * 100\% = 9.323 * 2 = 18.65\%$

Mass Percent as Conversion Factor

- The mass percent tell you the mass of a given element in 100g of the compound
- This can be used as a conversion factor to determine the gram of a given element in a sample of a compound

Example: Molar mass as Conversion Factor

- Find the mass of sodium in 5.6g of sodium chloride NaCl

Na: $\frac{5.6g}{58.44} * \frac{1 Na}{1 NaCl} * 22.99 g/mol$

$= 2.2g$ Na

Use Mol Mass of Na over NaCl mol mass instead... easier

Example: Mass Percent as a Conversion Factor

- Sodium chloride is 39% sodium by mass. Find the mass of sodium in 5.6g of sodium chloride

$5.6g\ NaCl * \frac{0.39 Na}{NaCl} = 2.2g$

Just multiply the grams of NaCl by the percent of Sodium to find the answer

Finding an Empirical formula

- Convert percent to grams
- Convert grams to moles
- Divide each number of moles by the smallest number of moles
- Multiply by a number to get a whole number ratio

Example

- Aspirin has a percent composition of 60.00% carbon, 4.48% hydrogen and 35.53% oxygen. What is the empirical formula for aspirin?

60g - C ÷ (12.01) = 4.996mol ÷ 2.221 = 2.250 * 4

4.48g - H (1.01) = 4.44mol ÷ 2.221 = 2 * 4

35.53 - O (16.00) = 2.221mol ÷ 2.221 = 1 * 4

$C_9H_8O_4$

> .9 round up

< .1 round down

Finding the Molecular Formula

- Determine the mass of the compound from it empirical formula
- Divide the empirical formula mass into the given mass of the molecule
- This should give a whole number
- Multiply the whole number by each of the subscripts

if C_2H_3O then

$C_2H_3O = 43.05 = \frac{86.03}{43.05} = 2 * C_2H_3O = C_4H_6O_2$

if C_2H_2O

Example: Molecular Formula

- Find the molecular formula for a compound that has a empirical formula of C_2H_3O and an experimental molar mass of 86.03 g/mol.

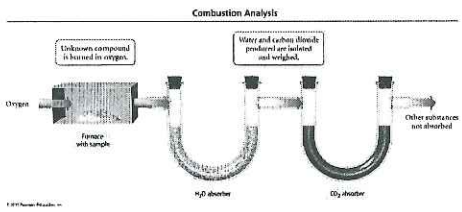
$C_2H_2O = 42.04\ g/mol \div 86.03\ g/mol$

$= 2.0463 = 2 * C_2H_2O$

$= C_4H_4O_2$

Ex: 119 pg 124
good practice

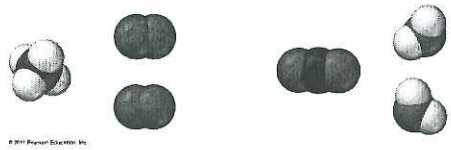
Combustion Analysis to Determine Empirical formula



Balancing

A chemical equation must be balanced for mass

$CH_4(g) + 2\,O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(g)$



Always reduce to the smallest whole # ratio of coeff

$2.445\,g$ of $CO_2 * \frac{mol}{44.01g} * \frac{1C}{1CO_2} = 0.05556\,mol\ C$

$0.6003\,g\ H_2O * \frac{mol}{18.02g} * \frac{2H}{1\,H_2O} = 0.06663\,mol\ H$

8/17/10

$0.8233g - (0.05556 * 12.01\,g/mol) - (0.06663 * 1.008\,g/mol)$

mol C mol H

$O = 0.0889\,g$

Example:Combustion Analysis to Determine Empirical formula

- The combustion of 0.8233g of a compound containing C, H and O produced 2.445g of carbon dioxide and 0.6003g of water. What is the empirical formula of the compound?

$C = \frac{0.0556}{0.00555} = 10$ $O = \frac{0.00555}{0.00555} = 1$

$H = \frac{0.06663}{0.00555} = 12$ $C_{10}H_{12}O$

$O = 0.0889\,g \div 16.00\,g/mol = 0.00555\,mol$

#29 & #91 for practice on CH 3

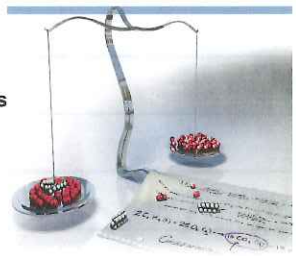
The End

#121 on CH 3 also

Lecture ends 32:50

Chapter 4

"Chemical Quantities and Aqueous Reactions"



Reaction Stoichiometry

- The coeffients in a chemical equation can be used to determine:
  - The amount of a second reagent needed
  - The amount of a product produced
- $2\ C_8H_{18}(l) + 25\ O_2(g) \rightarrow 16\ CO_2(g) + 18\ H_2O(g)$

$$2\ C_8H_{18} + 25\ O_2 \longrightarrow 16\ CO_2 + 18\ H_2O$$

Density gasoline = 0.8626 g/mL

Example

- Estimate the mass of CO_2 produced by the combustion of 4537 gal of gasoline.

$$4537\ gal\ C_8H_{18} * \frac{3785\ mL}{1\ gal} * \frac{0.8626\ g}{1\ mL} * \frac{mol\ C_8H_{18}}{114.3\ g} * \frac{16\ mol\ CO_2}{2\ mol\ C_8H_{18}} * \frac{44.01\ g}{mol\ CO_2}$$

$= 456288020.46$

$= 4.563 * 10^7\ g\ CO_2$

Limiting Reactant

- For reactions with multiple reactants, it is likely that one of the reactants will be completely used before the others
- The reactant that limits the amount of product is called the **limiting reactant**
- Reactants not completely consumed are called **excess reactants**
- The amount of product that can be made from the limiting reactant is called the **theoretical yield**

$$\frac{\text{actual yield}}{\text{Theo yield}} * 100\%$$

Percent Yield

- Many things can happen during the course of an experiment that cause the loss of product
- The amount of product that is made in a reaction is called the **actual yield**
- The efficiency of product recovery is generally given as the **percent yield**

Always start with Theo yield first

Example

- When 28.6 kg of C are allowed to react with 88.2 kg of TiO_2 in the reaction below, 42.8 kg of Ti are obtained. Find the Limiting Reactant, Theoretical Yield, and Percent Yield.

$$TiO_2(s) + 2\,C(s) \rightarrow Ti(s) + 2\,CO(g)$$

$$28.6\,Kg\,C * \frac{Kmol}{12.01\,Kg} * \frac{1\,Ti}{2\,C} * \frac{47.87\,Kg}{Kmol} = 57.0\,Kg\,Ti$$

Dont do unecessary conversions. You can use Kg/Kmol without having to change values

$$88.2\,Kg\,TiO_2 * \frac{Kmol}{79.87\,Kg} * \frac{1\,Ti}{1\,TiO_2} * \frac{47.87\,Kg}{Kmol} = 52.9\,Kg \Rightarrow \text{Limiting Reactant}$$

This is also my Theoretical Yield

$$\text{Percent Yield} = \frac{42.8\,Kg}{52.9\,Kg} * 100\% = 80.9\%$$

Practice

- How many grams of $N_2(g)$ can be made from 5.05 g of NH_3 reacting with 35.2 g of CuO?
  $2\ NH_3(g) + 3\ CuO(s) \rightarrow N_2(g) + 3\ Cu(s) + 3\ H_2O(l)$

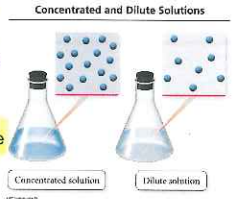
Solutions

- Homogeneous mixtures are called **solutions**
- The component of the solution that changes state is called the **solute**
- The component that keeps its state is called the **solvent**

Solvent typically used is water

Solution Concentration

- **Dilute solutions** have a small amount of solute compared to solvent
- **Concentrated solutions** have a large amount of solute compared to solvent
- Quantitatively, the relative amount of solute in the solution is called the **concentration**



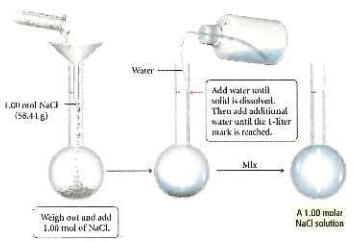
Solution Concentration
Molarity

- Moles of solute per 1 liter of solution
- Used because it describes how many molecules of solute in each liter of solution

$$M = \frac{mol}{L} = \frac{mmol}{mL}$$

Again no need to convert by a factor of one Thousand

Making a Solution



g → mol ÷ Volume (L) = M

Example

- Find the molarity of a solution that has 21.5 g KBr dissolved in 1.50 L of solution

$$\frac{21.5\ g\ KBr}{119.0\ g/mol} \div 1.50\ L = \boxed{0.120\ M}$$

$$0.1807\ mol * \frac{1}{1.50\ L} = \boxed{0.120\ M}$$

Using Molarity in Calculations

- Molarity shows the relationship between the moles of solute and liters of solution
- If a sugar solution concentration is 2.0 M, then 1 liter of solution contains 2.0 moles of sugar

Example

- How many liters of 0.245 M NaOH contains 0.176 mol NaOH?

$$0.176 \text{ mol NaOH} * \frac{\text{L}}{0.245 \text{ mol}}$$

$$= \boxed{.718 \text{ L}}$$

Dilution

- Often, solutions are stored as concentrated **stock solutions**
- To make solutions of lower concentrations from these stock solutions, more solvent is added
- The concentrations and volumes of the stock and new solutions are inversely proportional

$$M_1 \cdot V_1 = M_2 \cdot V_2$$

Know This

Example

- To what volume should you dilute 0.340 L of 14.0 M NaOH to make 2.50 M NaOH?

V_1 M_1 $=$

$(.340\,L)(14.0\,M) = (2.50\,M)(V_2)$

$V_2 = \frac{(.340\,L)(14.0\,M)}{(2.50\,M)} = \boxed{1.90\,L}$

The "of" means they go together V_1 & M_1

Solution Stoichiometry

- Since molarity relates the moles of solute to the liters of solution, it can be used to convert between amount of reactants and/or products in a chemical reaction

Volume A → Amount A (in moles) → Amount B (in moles) → Volume B

Volume of Solution just divide by Molarity

Volume of substance you'll be given density for calculation

Example

- What volume of 0.150 M KCl is required to completely react with 0.150 L of 0.175 M $Pb(NO_3)_2$ in the reaction?
  - $2\,KCl(aq) + Pb(NO_3)_2(aq) \rightarrow PbCl_2(s) + 2\,KNO_3(aq)$

$.150\,L\ Pb(NO_3)_2 * \frac{0.175\,mol}{L} * \frac{2\,KCl}{1\,Pb(NO_3)_2} * \frac{L}{0.150\,mol} = \boxed{.350\,L\ KCl}$

What Happens When a Solute Dissolves?

- There are attractive forces between the solute particles holding them together
- There are also attractive forces between the solvent molecules
- When we mix the solute with the solvent, there are attractive forces between the solute particles and the solvent molecules
- If the attractions between solute and solvent are strong enough, the solute will dissolve

Electrolytes and Nonelectrolytes

- Materials that dissolve in water to form a solution that will conduct electricity are called **electrolytes**
- Materials that dissolve in water to form a solution that will not conduct electricity are called **nonelectrolytes**

Anything that breaks up in water & have charged particles are considered electrolytes

Molecular Compounds have Covalent bonds & don't carry charge they are Non-electrolytes

Acids

- Acids are molecular compounds that **ionize** when they dissolve in water
- acids that ionize virtually 100% are called **strong acids**

$$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$

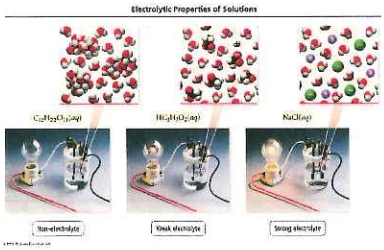
- acids that only ionize a small percentage are called **weak acids**

$$HF(aq) \Leftrightarrow H^+(aq) + F^-(aq)$$

Strong and Weak Electrolytes

- **Strong electrolytes** are materials that dissolve completely as ions
  - ionic compounds and strong acids → & strong bases
  - their solutions conduct electricity well
- **Weak electrolytes** are materials that dissolve mostly as molecules, but partially as ions
  - weak acids & weak bases
  - their solutions conduct electricity, but not well
- When compounds containing a polyatomic ion dissolve, the polyatomic ion stays together in water

Classes of Dissolved Materials



Solubility of Ionic Compounds

- Some ionic compounds, like NaCl, dissolve very well in water at room temperature
- Other ionic compounds, like AgCl, dissolve hardly at all in water at room temperature
- Compounds that dissolve in a solvent are said to be **soluble**, while those that do not are said to be **insoluble**

Solubility Rules
Compounds that Are Generally Soluble in Water

Compounds Containing the Following Ions are Generally Soluble	Exceptions (when combined with ions on the left the compound is insoluble)
Li^+, Na^+, K^+, NH_4^+	none
NO_3^-, $C_2H_3O_2^-$	none
Cl^-, Br^-, I^-	Ag^+, Hg_2^{2+}, Pb^{2+}
SO_4^{2-}	Ag^+, Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}

Solubility Rules
Compounds that Are Generally Insoluble

Compounds Containing the Following Ions are Generally Insoluble	Exceptions (when combined with ions on the left the compound is soluble or slightly soluble)
OH^-	Li^+, Na^+, K^+, NH_4^+, Ca^{2+}, Sr^{2+}, Ba^{2+}
S^{2-}	Li^+, Na^+, K^+, NH_4^+, Ca^{2+}, Sr^{2+}, Ba^{2+}
CO_3^{2-}, PO_4^{3-}	Li^+, Na^+, K^+, NH_4^+

Precipitation Reactions

- Reactions between aqueous solutions of ionic compounds that produce an ionic compound that is insoluble in water are called **precipitation reactions** and the insoluble product is called a **precipitate**

Precipitate Formation = Reaction

No Precipitate Formation = No Reaction

$KI(aq) + NaCl(aq) \rightarrow KCl(aq) + NaI(aq)$
all ions still present, ∴ no reaction

Process for Predicting the Products of a Precipitation Reaction

- Determine what ions each aqueous reactant has
- Determine formulas of possible products
- Determine Solubility of Each Product in Water
- If neither product will precipitate, write **no reaction** after the arrow

Process for Predicting the Products of a Precipitation Reaction

If either product is insoluble, write the formulas for the products after the arrow – writing (*s*) after the product that is insoluble and will precipitate, and (*aq*) after products that are soluble and will not precipitate

Balance the equation

K^+ CO_3^{2-} Ni^{2+} Cl^-

Example

- Write the equation for the precipitation reaction between an aqueous solution of potassium carbonate and an aqueous solution of nickel(II) chloride

$K_2CO_3(aq) + NiCl_2(aq) \rightarrow NiCO_3(s) + KCl(aq)$

$K_2CO_3(aq) + NiCl_2(aq) \rightarrow NiCO_3(s) + 2\,KCl(aq)$

if your Double displacement equation is not balanzing nicely look at your internal balancing because it will be off Double Displacement eq. balance very easily

Ionic Equations

- Consider the following equation

$2\,KOH(aq) + Mg(NO_3)_2(aq) \longrightarrow 2\,KNO_3(aq) + Mg(OH)_2(s)$

- The complete ionic equation can be generated by looking at the dissolved species

Spectator ions have to be the same charge & in the same amount

Exam will have a Redox Calculation

Acid-Base Reactions

- Also called **neutralization reactions** because the acid and base neutralize each other's properties

Acids and Bases in Solution

- Acids ionize in water to form H^+ ions
  - more precisely, the H from the acid molecule is donated to a water molecule to form **hydronium ion, H_3O^+**
- Bases dissociate in water to form OH^- ions
  - bases, like NH_3, that do not contain OH^- ions, produce OH^- by pulling H off water molecules

Acid + Water typically gives us H_3O^+

Strong Acids on Test
HCl, H_2SO_4, HNO_3

Acids and Bases in Solution

- In the reaction of an acid with a base, the H^+ from the acid combines with the OH^- from the base to make water
- The cation from the base combines with the anion from the acid to make the salt

acid + base → salt + water

Strong Bases on Test
$NaOH$, KOH

Exam Question

Example

- Write the balanced equation then the ionic, and net-ionic equation for the reaction of aqueous nitric acid with aqueous calcium hydroxide

$2HNO_3(aq) + Ca(OH)_2(aq) \rightarrow 2H_2O(l) + Ca(NO_3)_2(aq)$

Balance everything but hydrogen. Leave it for last

Tip: Number of Hydrogens on Acids must equal the number of bases (Hydroxides) for reactants

Titration

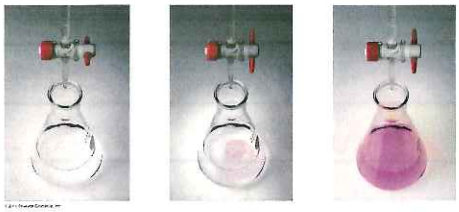
- Often in the lab, a solution's concentration is determined by reacting it with another material and using stoichiometry – this process is called **titration**
- In the titration, the unknown solution is added to a known amount of another reactant until the reaction is just completed, at this point, called the **endpoint**, the reactants are in their stoichiometric ratio

Acid-Base Titrations

- In acid-base titrations, because both the reactant and product solutions are colorless, a chemical is added that changes color when the solution undergoes large changes in acidity/alkalinity
  - the chemical is called an **indicator**
- at the endpoint of an acid-base titration, the number of moles of H^+ equals the number of moles of OH^-
  - aka the **equivalence point**

Titration

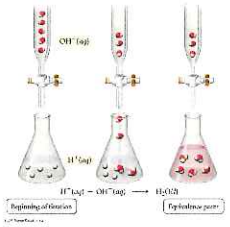
Indicator in Titration



The lighter the pink the closer to neutrality

Titration

Acid–Base Titration



$HCl + NaOH \longrightarrow NaCl + H_2O$

Example

- The titration of 10.00 mL of HCl solution of unknown concentration requires 14.67 mL of 0.100 M NaOH solution to reach the end point. What is the concentration of the unknown HCl solution?

$$14.67\ mL\ NaOH * \frac{0.100\ mmol}{mL} * \frac{1\ HCL}{1\ NaOH} * \frac{1}{10.00\ mL} = 0.1467\ M = \boxed{0.147\ M}$$

$$M = \frac{mol}{L} = \frac{mmol}{mL}$$

On Exam

Gas Evolving Reactions

- Some reactions form a gas directly from the ion exchange
  $K_2S(aq) + H_2SO_4(aq) \rightarrow K_2SO_4(aq) + H_2S(g)$
- Other reactions form a gas by the decomposition of one of the ion exchange products into a gas and water
  $K_2SO_3(aq) + H_2SO_4(aq) \rightarrow K_2SO_4(aq) + H_2SO_3(aq)$
  $H_2SO_3 \rightarrow H_2O(l) + SO_2(g)$

H_2CO_3 & H_2SO_3 are unstable intermediates since they do not become H_2CO_3 or H_2SO_3. They become $H_2O(L) + CO_2(G)$ & $H_2O(L) + SO_2(G)$

$Na_2CO_3 + HNO_3 \rightarrow NaNO_{3(aq)} + H_2O(L) + CO_2(G)$

Example

- When an aqueous solution of sodium carbonate is added to an aqueous solution of nitric acid, a gas evolves

$Na_2CO_3 + 2HNO_3 \rightarrow 2NaNO_{3(aq)} + H_2O(L) + CO_2(G)$

Redox

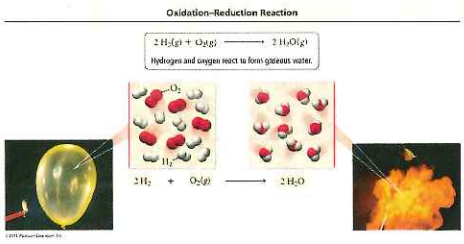
- Reactions involve transferring electrons from one atom to another are called **oxidation-reduction reactions**

Redox rxns always involve electrons

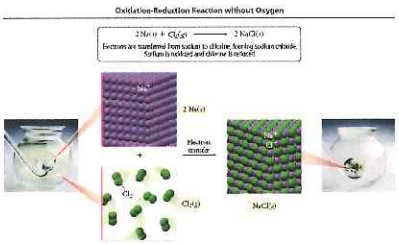
All non-nuclear explosions are redox rxn

Electrons moving very fast goes BOOM!

Combustion as Redox



Redox without Combustion



Oxidation and Reduction

- reactions where electrons are transferred from one atom to another are redox reactions
- atoms that lose electrons are being oxidized, atoms that gain electrons are being reduced

Electron Bookkeeping

- chemists assign a number to each element in a reaction called an **oxidation state** that allows them to determine the electron flow in the reaction
  - even though they look like them, **oxidation states are not ion charges!**
    - oxidation states are imaginary charges assigned based on a set of rules
    - ion charges are real, measurable charges

Oxidation States are imaginary

Exam

Rules for Assigning Oxidation States

- rules are in order of priority

1. free elements have an oxidation state = 0
   - Na = 0 and Cl_2 = 0 in 2 Na(*s*) + Cl_2(*g*)
2. monatomic ions have an oxidation state equal to their charge
   - Na = +1 and Cl = -1 in NaCl
3. (a) the sum of the oxidation states of all the atoms in a compound is 0
   - Na = +1 and Cl = -1 in NaCl, (+1) + (-1) = 0

Need to Know these steps or how to do this for the exam

Rules for Assigning Oxidation States

3. (b) the sum of the oxidation states of all the atoms in a polyatomic ion equals the charge on the ion
   - N = +5 and O = -2 in NO_3^-, (+5) + 3(-2) = -1
5. (a) Group I metals have an oxidation state of +1 in all their compounds
   - Na = +1 in NaCl
6. (b) Group II metals have an oxidation state of +2 in all their compounds
   - Mg = +2 in $MgCl_2$

Group I & II metals have charge priority

Rules for Assigning Oxidation States

5. In their compounds, nonmetals have oxidation states according to the following table below

Exam

Rules for Assigning Oxidation States

Nonmetal	Oxidation State	Example
F	-1	CF_4
H	+1	CH_4
O	-2	CO_2
Group 7A	-1	CCl_4
Group 6A	-2	CS_2
Group 5A	-3	NH_3

This is on page 164

Priority after metals starting from F down

Practice – Assign an Oxidation State to Each Element in the following

Br_2 – 0

K^+ – +1

LiF – Li = +1 so F = -1

CO_2 – C = +4 O = -2

SO_4^{-2} – S = +6 O = -2

Na_2O_2 – Na = +1 O = -1

Free elements have an Oxidation state of Zero = 0

On exam = reduction half, oxidation half, oxidising reducing Agent

$$I^-_{(aq)} + HNO_{3(aq)} \longrightarrow I_{2(s)} + NO_{(g)} + H_2O_{(l)}$$

(in Acid) H_3O^+

$$\underline{I^-} \longrightarrow I_2 + 2e^-$$ this oxidised

$$[e^- + H^+ \longrightarrow H_2]\times 3$$ this reduced

$$3e^- + 3H^+ \longrightarrow 3H_2$$

$$6I^- + 2HNO_3 \longrightarrow I_2 + 2NO$$

1 oxygen

$$6I^- + 2HNO_3 \longrightarrow I_2 + 2NO + 2H_2O$$

4 Now Hydrogen is missing so add H^+ to balance

$$4H^+ + 6I^- + 2HNO_3 \longrightarrow I_2 + 2NO + 2H_2O$$

3 Now make sure both sides balance with charge

$$H^+ + H_2O \rightarrow H_3O^+$$

4 Now add H_2O to both sides to make up for the H^+ we added

$$4H^+ + 6I^- + 2HNO_3 \longrightarrow I_2 + 2NO + 2H_2O$$

$$4H_3O^+_{(aq)} + 6I^-_{(aq)} + 2HNO_{3(aq)} \longrightarrow I_{2(s)} + 2NO_{(g)} + 6H_2O_{(l)}$$

on exam = reduction half, oxidation half, oxidizing reducing Agent

on Exam

Pg 858

* Gases, Solids & Liquids have a Zero charge

40) a.

$$\underset{-1}{I^-}(aq) + \underset{+3\ -4}{NO_2^-}(aq) \longrightarrow \overset{0}{I_2}(s) + \underset{+2\ -2}{NO}(g) \quad (\text{in Acid})$$

H_2O, H_3O^+

$$\overset{-2}{2I^-} \longrightarrow \overset{0}{I_2} + 2e^- \quad \text{this oxidized}$$

$$[e^- + N^{+3} \longrightarrow N^{+2}] * 2 \quad \text{this reduced}$$

$$2e^- + 2N^{+3} \longrightarrow 2N^{+2}$$

$$2I^{-1} + \underbrace{2NO_2^-}_{4\ \text{oxygen}} \longrightarrow I_2 + \underbrace{2NO}_{2\ \text{Oxygen}}$$

• 1 + H_2O for each missing Oxygen so...

$$2I^{-1} + 2NO_2^- \longrightarrow I_2 + 2NO + 2H_2O$$

• 2 Now Hydrogen is missing so add H^+ to balance

$$4H^+ + 2I^{-1} + 2NO_2^- \longrightarrow I_2 + 2NO + 2H_2O$$

• 3 No make sure both sides balance with charge

$$H^+ + H_2O \longrightarrow H_3O^+$$

• 4 Now add H_2O to both sides to make up for the H^+ we added

$$\underbrace{4H^+}_{\text{So } +4H_2O} + 2I^- + 2NO_2^- \longrightarrow I_2 + 2NO + \underbrace{2H_2O}_{+4H_2O}$$

$$4H_3O^+(aq) + 2I^-(aq) + 2NO_2^-(aq) \longrightarrow I_2(s) + 2NO(g) + 6H_2O(l)$$

Oxidation = oxidation state increased
Reduction = oxidation state decreases
Oxidation = $Na^0 \longrightarrow Na^+$
Reduction = $O_2 \longrightarrow 2O^{-2}$

8/26/10

Examples Redox

Combustion Reactions

- Reactions in which $O_2(g)$ is a reactant are called **combustion reactions**

Combustion Products

Reactant	Combustion Product
contains C	$CO_2(g)$
contains H	$H_2O(g)$
contains S	$SO_2(g)$
contains N	$NO(g)$ or $NO_2(g)$
contains metal	$M_2O_n(s)$

Read pg 816 – 819 to understand better if needed

oxidation

$$\overset{-4}{C}\overset{+1}{H_4} + 2\overset{0}{O_2} \longrightarrow \overset{+4}{C}\overset{-2}{O_2} + 2\overset{+1}{H_2}\overset{-2}{O}$$

Reduction

Metal Redox = $Zn^{0}_{(s)} + 2Ag^{+1}_{(aq)} \longrightarrow Zn^{+2}_{(aq)} + 2Ag^{0}_{(s)}$

8/26/10

$Zn \longrightarrow Zn + 2e^-$ $\quad 2(1e^- + Ag^+ \longrightarrow Ag^0)$

to balance

$(2e^- + 2Ag^+ \longrightarrow 2Ag^0)$

Examples

- Combustion of $C_3H_7OH(l)$
- Combustion of $CH_3NH_2(g)$

* Pick a problem from 858 #39 & #41

The End

- Oxidation & Reduction <u>must</u> occur simontaniously
- Reducing Agent = undergoes oxidation
- Oxidizing Agent = undergoes reduction

Oxygen is an Oxidizing Agent because it reduces everybody by giving e^- & reducing the overall charge

Acid + Base Redox

a) $\overset{-2}{PbO_2}{}_{(s)} + \overset{-1}{I^-}{}_{(aq)} \longrightarrow Pb^{2+}{}_{(aq)} + \overset{0}{I_2}{}_{(s)}$

$+4 \quad -4$

$2e^- + Pb^{+4} \longrightarrow Pb^{+2}$ $\qquad$ $2I^- \longrightarrow I_2^0 + 2e^-$

$PbO_2 + 2I^- \longrightarrow Pb^{+2} + I_2 + 2H_2O$

$4H^+ + PbO_2 + 2I^- \longrightarrow Pb^{+2} + I_2 + 2H_2O$

Acid $\quad 4H_3O^+ + PbO_2 + 2I^- \longrightarrow Pb^{+2} + I_2 + 6H_2O$

Base $\quad \underbrace{4H^+}_{+4OH^-} + PbO_2 + 2I^- \longrightarrow Pb^{+2} + I_2 + 2H_2O \underbrace{\quad}_{+4OH^-}$

$\underbrace{4H_2O_{(L)}}_{\text{cancel}} + PbO_2 + 2I^-_{(aq)} \longrightarrow Pb^{+2}_{(aq)} + I_{2(s)} + \underbrace{2H_2O}_{\text{cancel}} + 4OH^-$

$2H_2O_{(L)} + PbO_{2(s)} + 2I^-_{(aq)} \longrightarrow Pb^{+2}_{(aq)} + I_{2(s)} + 4OH^-_{(aq)}$

$$PbO_{2(s)} + I^-_{(aq)} \longrightarrow Pb^{2+}_{(aq)} + I_{2(s)}$$

$$2e^- + Pb^{4+} \longrightarrow Pb^{2+} \qquad 2I^- \longrightarrow I_2 + 2e^-$$

$$PbO_2 + 2I^- \longrightarrow Pb^{2+} + I_2 + 2H_2O$$

$$4H^+ + PbO_2 + 2I^- \longrightarrow Pb^{2+} + I_2 + 2H_2O$$

$$4H_2O + PbO_2 + 2I^- \longrightarrow Pb^{2+} + I_2 + 6H_2O$$

$$4H^+ + PbO_2 + 2I^- \longrightarrow Pb^{2+} + I_2 + 2H_2O$$

$$PbO_{2(s)} + 2I^-_{(aq)} \longrightarrow Pb^{2+}_{(aq)} + I_{2(s)} + 2H_2O + 4OH^-_{(aq)}$$

$$2H_2O_{(l)} + PbO_{2(s)} + 2I^-_{(aq)} \longrightarrow Pb^{2+}_{(aq)} + I_{2(s)} + 4OH^-_{(aq)}$$

Pg 858

Balance Acid Redox Reactions

#39

a) $\overset{+4}{Pb}\overset{-2}{O_2}(s) + \overset{-1}{I^-}(aq) \longrightarrow \overset{2+}{Pb^{2+}}(aq) + \overset{0}{I_2}(s)$

$2e^- + \overset{+4}{Pb} \longrightarrow Pb^{2+}$ $\qquad$ $2\overset{-2}{I^-} \longrightarrow I_2^0 + 2e^-$

$PbO_2(s) + 2I^-(aq) \longrightarrow Pb^{2+}(aq) + I_2(s)$

so $+2H_2O$

$PbO_2(s) + 2I^-(aq) \longrightarrow Pb^{2+}(aq) + I_2(s) + 2H_2O(L)$

so add H^+

$4H^+ + 2PbO_2(s) + 2I^-(aq) \longrightarrow Pb^{2+}(aq) + I_2(s) + 2H_2O(L)$

$$\boxed{4H_3O^+(aq) + 2I^-(aq) \longrightarrow Pb^{2+}(aq) + I_2(s) + 6H_2O(L)}$$

b) $\overset{+4}{S}\overset{-6}{O_3^{2-}}(aq) + \overset{+7}{Mn}\overset{-8}{O_4^-}(aq) \longrightarrow \overset{+6}{S}\overset{-8}{O_4^{2-}}(aq) + \overset{+2}{Mn^{2+}}(aq)$

$\underset{+4}{S} \longrightarrow \underset{+6}{S} + 2e^-$ $\qquad$ $5e^- + \underset{+7}{Mn} \longrightarrow \underset{+2}{Mn}$

$5S^{+4} \longrightarrow 5S^{+6} + 10e^-$ $\qquad$ $10e^- + 2Mn^{+7} \longrightarrow 2Mn^{+2}$

$5SO_3^{2-} + 2MnO_4^- \longrightarrow 5SO_4^{2-} + 2Mn^{2+} + 3H_2O$

$6H^+(aq) + 5SO_3^{2-}(aq) + 2MnO_4^-(aq) \longrightarrow 5SO_4^{2-}(aq) + 2Mn^{2+}(aq) + 3H_2O(L)$

$$\boxed{6H_3O^+(aq) + 5SO_3^{2-}(aq) + 2MnO_4^-(aq) \longrightarrow 5SO_4^{2-}(aq) + 2Mn^{2+}(aq) + 9H_2O(L)}$$

c) $\underset{+4}{S_2}\underset{-6}{O_3^{2-}}{}_{(aq)} + \underset{0}{Cl_2}{}_{(g)} \longrightarrow \underset{+6}{S}\underset{-8}{O_4^{2-}}{}_{(aq)} + \underset{-1}{Cl^-}{}_{(aq)}$

$S_2^{+4} \longrightarrow 2S^{+6} + 8e^-$ $\qquad$ $2e^- + Cl_2^0 \longrightarrow 2Cl^-$

$8e^- + 4Cl_2 \longrightarrow 8Cl^-$

$S_2O_3^{2-} + 4Cl_2 \longrightarrow 2SO_4^{2-} + 8Cl^-$

$S_2O_3^{2-}{}_{(aq)} + 4Cl_{2(G)} + 5H_2O_{(L)} \longrightarrow 2SO_4^{2-}{}_{(aq)} + 8Cl^-{}_{(aq)} + 10H^+$

$$\boxed{S_2O_3^{2-}{}_{(aq)} + 4Cl_{2(G)} + 15H_2O_{(L)} \longrightarrow 2SO_4^{2-}{}_{(aq)} + 8Cl^-{}_{(aq)} + 10H_3O^+{}_{(aq)}}$$

#41 a) $H_2O_{2(aq)} + ClO_{2\,(aq)} \longrightarrow ClO_2^-{}_{(aq)} + O_{2\,(G)}$

$Cl^{+4} \longrightarrow Cl^{+3} + e^-$ $\qquad$ $O_2^{-1} \longrightarrow O_2^0 + 2e^-$

$2Cl^{+4} \longrightarrow 2Cl^{+3} + 2e^-$

$\underbrace{H_2O_2}_{+2OH^-} + 2ClO_2 \longrightarrow 2ClO_2^- + O_2 \underbrace{\quad}_{+2OH^-}$

$+2OH^- + H_2O_2 + 2ClO_2 \longrightarrow 2ClO_2^- + O_2 + 2OH^-$

#40 a) $I^-_{(aq)} + NO_2^-{}_{(aq)} \rightarrow I_{2(s)} + NO_{(G)}$

$2I^-_{(aq)} \rightarrow I_2^0 + 2e^-$ $\quad$ $e^- + N^{+3} \rightarrow N^{+2}$

$2e^- + 2N^{+3} \rightarrow 2N^{+2}$

$2I^-_{(aq)} + 2NO_2^-{}_{(aq)} \rightarrow I_{2(s)} + 2NO_{(G)} + 2H_2O_{(L)}$

$4H^+_{(aq)} + 2I^-_{(aq)} + 2NO_2^-{}_{(aq)} \rightarrow I_{2(s)} + 2NO_{(G)} + 2H_2O_{(L)}$

$$\boxed{4H_3O^+_{(aq)} + 2I^-_{(aq)} + 2NO_2^-{}_{(aq)} \rightarrow I_{2(s)} + 2NO_{(G)} + 6H_2O_{(L)}}$$

b) $ClO_4^-{}_{(aq)} + Cl^-_{(aq)} \rightarrow ClO_3^-{}_{(aq)} + Cl_{2(G)}$

$2e^- + Cl^{+7} \rightarrow Cl^{+5}$ $\quad$ $2Cl^- \rightarrow Cl_2^0 + 2e^-$

$ClO_4^- + 2Cl^- \rightarrow ClO_3^- + Cl_2(G) + H_2O_{(L)}$

$2H^+ + ClO_4^- + 2Cl^- \rightarrow ClO_3^- + Cl_2 + H_2O_{(L)}$

? $$\boxed{2H_3O^+_{(aq)} + ClO_4^-{}_{(aq)} + 2Cl^-_{(aq)} \rightarrow ClO_3^-{}_{(aq)} + Cl_{2(G)} + 3H_2O_{(L)}}$$

c) $NO_{3\,(aq)}^{-} + Sn^{2+}_{(aq)} \longrightarrow Sn^{4+}_{(aq)} + NO_{(G)}$

$3e^- + N^{+5} \longrightarrow N^{2+}$ $\qquad$ $Sn^{2+} \longrightarrow Sn^{4+} + 2e^-$

$6e^- + 2N^{+5} \longrightarrow 2N^{2+}$ $\qquad$ $3Sn^{2+} \longrightarrow 3Sn^{4+} + 6e^-$

$2NO_{3\,(aq)}^{-} + 3Sn^{2+}_{(aq)} \longrightarrow 3Sn^{4+}_{(aq)} + 2NO_{(G)}$

$2NO_{3\,(aq)}^{-} + 3Sn^{2+}_{(aq)} \longrightarrow 3Sn^{4+}_{(aq)} + 2NO_{(G)} + 4H_2O_{(L)}$

$8H^{+}_{(aq)} + 2NO_{3\,(aq)}^{-} + 3Sn^{2+}_{(aq)} \longrightarrow 3Sn^{4+}_{(aq)} + 2NO_{(G)} + 4H_2O_{(L)}$

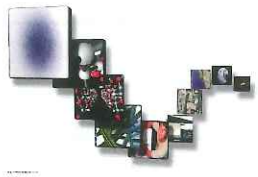
? $\boxed{8H_3O^{+}_{(aq)} + 2NO_{3\,(aq)}^{-} + 3Sn^{2+}_{(aq)} \longrightarrow 3Sn^{4+}_{(aq)} + 2NO_{(G)} + 12H_2O_{(L)}}$

1s 2s 2p 3s 3p
4s 3d 4p
5s 4d 5p
6s 4f 5d 6p
7s 5f 6d 7p

Know all This Exam

Chapter 7

"The Quantum-Mechanical Model of the Atom"



Shell: 1, 2, 3 (n)
Subshell: s, p, d, f (l)

orbital: — (m_l)
spin: ↑ (m_s)

Very Small

- Electrons are incredibly small
- Electron behavior determines much of the behavior of atoms (reactivity)
- Directly observing electrons in the atom is not possible yet, the electron is so small that observing it may change its behavior

A Theory

- The **quantum-mechanical model** explains the manner electrons exist and behave in atoms
- Helps us understand and predict the properties of atoms that are directly related to the behavior of the electrons

$n = 1, 2, 3, 4$

$l = s, p, d, f$
$\quad\;\; 0, 1, 2, 3$

$m_l =$ s: 0; p: −1, 0, +1; d: −2, −1, 0, 1, 2

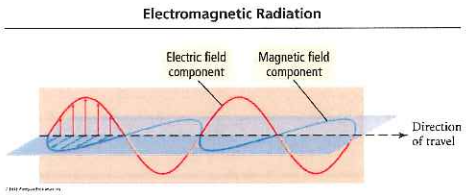
$m_s =$ ↑ $= \frac{1}{2}$ ↓ $= -\frac{1}{2}$

Light

- Light is a form of **electromagnetic radiation**
  - composed of perpendicular oscillating waves, one for the electric field and one for the magnetic field
    - an electric field is a region where an electrically charged particle experiences a force
    - a magnetic field is a region where an magnetized particle experiences a force
- All electromagnetic waves move through space at the same, constant speed
  3.00×10^8 m/s in a vacuum = **the speed of light, c**

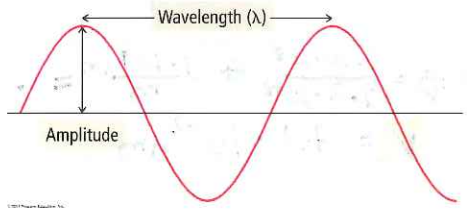
Sound Speed v.s. Light Speed

Electromagnetic Radiation



oscilates between electric & Magnetic field

Wave Characteristics



Characterizing Waves

- The **amplitude** is the height of the wave
  - the amplitude is a measure of how intense the light is
- The **wavelength, (λ)** is a measure of the distance covered by the wave
- The distance from one crest to the next the **frequency, (ν)** is the number of waves that pass a point in a given period of time
  - the number of waves = number of cycles

# of cycles of waves

as wavelength increases frequency does what?

a) increase
b) decrease ✓
c) nothing

The Relationship Between Wavelength and Frequency

- For waves traveling at the same speed, the shorter the wavelength, the more frequently they pass
- This means that the wavelength and frequency of electromagnetic waves are inversely proportional
- Note: $c = \lambda * \nu$

$c = m/s$

$\nu = Hz = 1/s = s^{-1}$

$\lambda = m$

$c = \nu \cdot \lambda$

Example

- Calculate the wavelength of red light with a frequency of $4.60 \times 10^{14}\ s^{-1}$

$\lambda = \frac{c}{\nu}$ $\quad \lambda = \frac{3.00 \cdot 10^{8} m/s}{4.60 \cdot 10^{14} s^{-1}}$

$\lambda = 6.521739 \cdot 10^{-7} * (1 \cdot 10^{9})$

$\lambda = 652 nm$

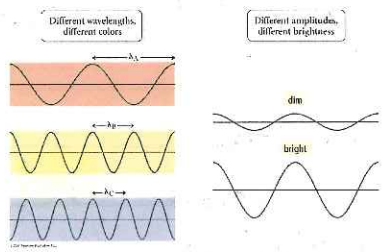
$\frac{1 \cdot 10^{9} nm}{m}$

Quiz: Determine wavelength in nm of frequency $4.2 \cdot 10^{18} Hz$

$\lambda = \frac{c}{\nu} = \frac{3.00 \cdot 10^{8}}{4.2 \cdot 10^{18}} * \frac{1 \cdot 10^{9} nm}{m} = .071429 nm$

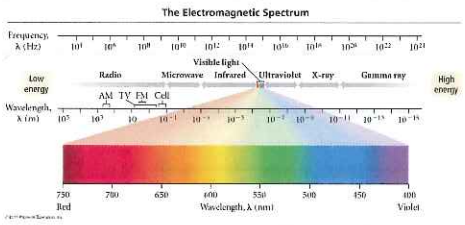
✓ $= 7.1 \cdot 10^{-2} nm$

Amplitude & Wavelength



Electromagnetic Radiation: What We See

Electromagnetic Spectrum



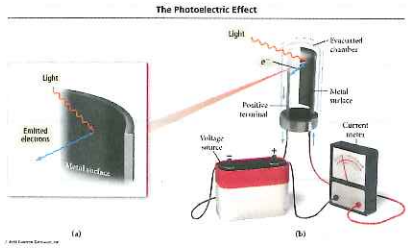
Know the order of waves

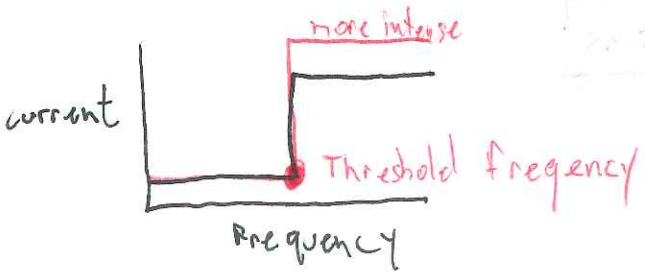
Radio — Lowest
Microwave
Infrafred
Visible light
Ultra violet
X-ray
Gama ray — Highest

The Electromagnetic Spectrum

- Visible light comprises only a small fraction of all the wavelengths of light – called the **electromagnetic spectrum**
- Short wavelength (high frequency) light has high energy

The Photoelectric Effect





The Photoelectric Effect
The Problem

- In experiments with the photoelectric effect, it was observed that there was a maximum wavelength for electrons to be emitted
  - called the **threshold frequency**
  - regardless of the intensity
- It was also observed that high frequency light with a dim source caused electron emission without any lag time

Explanation

- It was proposed that the light energy was delivered to the atoms in packets, called **quanta** or **photons**
- The energy of a photon of light was directly proportional to its frequency
  - the proportionality constant is called **Planck's Constant, (h)** and has the value 6.626×10^{-34} J·s
- $E_{photon} = h\nu = hc/\lambda$

Photons are energy

Example

- Calculate the number of photons in a laser pulse with wavelength 435 nm and total energy 3.23 mJ

$$E_{photon} = \frac{(6.626 \cdot 10^{-34}\ J \cdot s)(3.00 \cdot 10^{8}\ m/s)}{(4.35 \cdot 10^{-7}\ m)}$$

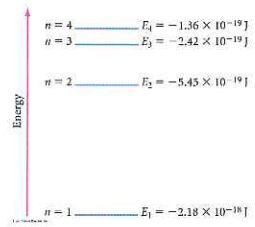
$$E_{photon} = 4.57 \cdot 10^{-19}\ J/photon$$

$$3.23\ mJ * \frac{J}{1000\ mJ} * \frac{photon}{4.57 \cdot 10^{-19}\ J}$$

$$= \boxed{7.07 * 10^{15}\ photons}$$

Principal Energy Levels in Hydrogen

En = -2.18 x 10^{-18} J (1/n^2)

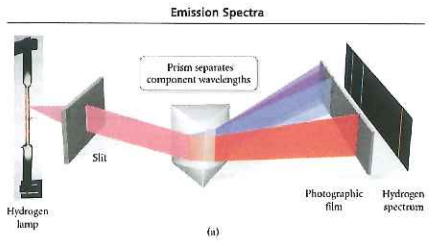


we can only see energy emitted in shell 2

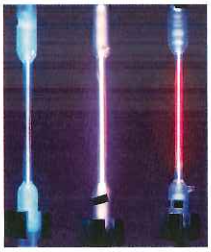
Spectra

- When atoms or molecules absorb energy, that energy is often released as light energy
  - fireworks, neon lights, etc.
- When that light is passed through a prism, a pattern is seen that is unique to that type of atom or molecule – the pattern is called an **emission spectrum**
- Rydberg analyzed the spectrum of hydrogen and found that it could be described with an equation that involved an inverse square of integers

Emission Spectra



Exciting Gas Atoms to Emit Light with Electrical Energy



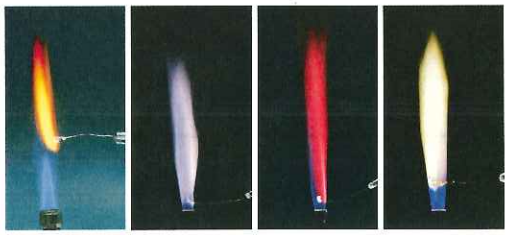
Hg He H

Examples of Spectra

Helium spectrum

Barium spectrum

Identifying Elements with Flame Tests

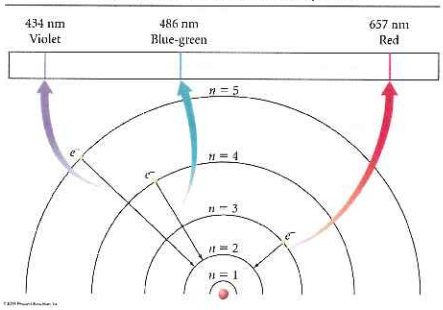


Na K Li Ba

Bohr's Model

- Neils Bohr proposed that the electrons could only have very specific amounts of energy
  - fixed amounts = **quantized**
- The electrons traveled in orbits that were a fixed distance from the nucleus
  - **stationary states**
- Electrons emitted radiation when they "Fell" from an orbit with higher energy down to an orbit with lower energy

The Bohr Model and Emission Spectra

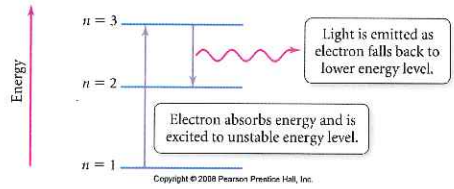


* fall to energy level 3 = infrared

* fall to energy level 1 = ultra violet

Electron Transitions

Excitation and Radiation



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Predicting the Spectrum of Hydrogen

- The wavelengths of lines in the emission spectrum of hydrogen can be predicted by calculating the difference in energy between any two states
- For an electron in energy state *n*, there are (*n* – 1) energy states it can transition to, therefore (*n* – 1) lines it can generate
- Both the Bohr and Quantum Mechanical Models can predict these lines very accurately

Predicting the Spectrum of Hydrogen

$$E_{\text{photon released}} = -\Delta E_{\text{hydrogen electron}} = -(E_{final} - E_{initial})$$

$$h\nu = \frac{hc}{\lambda} = -\left[-2.18\times10^{-18}\ \text{J}\left(\frac{1}{n_{final}^{2}}\right) - \left(-2.18\times10^{-18}\ \text{J}\left(\frac{1}{n_{initial}^{2}}\right)\right)\right]$$

$1/\lambda = R\,((1/2^2) - (1/n^2))$ $\quad R = 1.0974 \times 10^{7}\ m^{-1}$

Calculating the change in Energy

- The equation:
- $\Delta E = E_{final} - E_{initial}$
- $\Delta E = -Rhc((1/n^2_{final}) - (1/n^2_{initial}))$
- $-RHc = -2.179 \times 10^{-18}$ J/atom
- $-Rhc = -1312$ kJ/mole

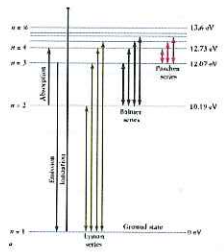
Hydrogen Energy Transitions

- Two Series
  - Lyman Series
    - Final level is n=1
  - Balmar Series
    - Final level is n=2

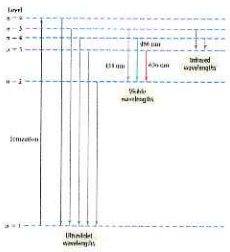
Lyman ends in $n = 1$

Balmer ends in $n = 2$

Hydrogen Energy Transitions



Hydrogen Energy Transitions



$R = 2.18 \cdot 10^{-18} J$

Example

- Calculate the change in energy for an electron that falls from n=4 to n=2

$E = 2.18 * 10^{-18} J \left(\frac{1}{n^2} - \frac{1}{n^2}\right)$

$E = 2.18 * 10^{-18} J \left(\frac{1}{4} - \frac{1}{16}\right)$

$\boxed{E = 4.09 * 10^{-19} J}$

* $*10^{-19} J$ is visible light

Calculating wavelength from change in energy

- Recall:
  - $E_{photon} = h\nu = hc/\Lambda$
  - $\lambda = hc/E_{photon}$

$\lambda = \frac{(6.626 \cdot 10^{-34} J \cdot s)(3.00 * 10^{8} m/s)}{}$

Relationship between wavelength and energy levels

- First proposed by Johann Balmer
- Later by Johannes Rydberg
- Rydberg Equation:

n_f

$1/\lambda = R((1/2^2) - (1/n^2))$ when n>2

~~R = 1.0974 x 10^-7 m^-1~~

$R = 1.0974 * 10^{7} m^{-1}$

Quiz: Determine wavelength (nm) from $n=3$ to $n=5$

$E = R\left(\frac{1}{9} - \frac{1}{25}\right) \longrightarrow E = 1.0974 * 10^{7}\left(\frac{16}{225}\right)$

$E = 780373^{-1}$ $\quad \lambda = 1.28 * 10^{-6} * \frac{1 \cdot 10^{9}\,nm}{m}$

$\boxed{\lambda = 1280\,nm}$

9/7/10

Example

- Calculate the wavelength of light emitted when the hydrogen electron transitions from $n = 4$ to $n = 2$

$\frac{1}{\lambda} = 1.097 * 10^{7} m^{-1}\left(\frac{1}{4} - \frac{1}{16}\right)$

$\frac{1}{\lambda} = 1.097 * 10^{7} m^{-1}\left(\frac{3}{16}\right)$

$\frac{1}{\lambda} = 2056875\,m^{-1}$

$\lambda = 4.862 * 10^{-7} m * \frac{1 \cdot 10^{9}\,nm}{m}$

$\lambda = 486.2\,nm$

Wave Behavior of Electrons

- De Broglie proposed that particles could have wave-like character
- Because it is so small, the wave character of electrons is significant
- De Broglie predicted that the wavelength of a particle was inversely proportional to its momentum
- $\lambda = h/mv$ *(v is velocity)*

* Quiz: As λ increases energy also ~~[illegible]~~ decreases

$\lambda = \frac{h}{mv}$ $\quad \frac{J \cdot s \cdot s}{\frac{m}{s} \cdot Kg} =$

Example

- Calculate the wavelength of an electron traveling at 2.65×10^{6} m/s

$\lambda = \frac{6.626 * 10^{-34}\,J \cdot s}{(2.65 * 10^{6} m/s)(9.109 * 10^{-31} Kg)}$

$\lambda = 2.74 * 10^{-10} m * \frac{1 \cdot 10^{9}\,nm}{m}$

$\boxed{\lambda = 0.274\,nm}$

$J = Kg\,m^{2}/s^{2}$

$\frac{Kg\,m^{2}/s^{2} \cdot s^{2}}{m \cdot Kg} = m$

Complimentary Properties

- When you try to observe the wave nature of the electron, you cannot observe its particle nature – and visa versa
- The wave and particle nature of nature of the electron are **complimentary properties**
  - as you know more about one you know less about the other

Uncertainty Principle

- Heisenberg stated that the product of the uncertainties in both the position and speed of a particle was inversely proportional to its mass
- The means that the more accurately you know the position of a small particle, like an electron, the less you know about its speed
  - and visa-versa

Make a choice
Do you want to know where it is or how fast it is going?

Determinacy vs. Indeterminacy

- According to classical physics, particles move in a path **determined** by the particle's velocity, position, and forces acting on it
  - determinacy = definite, predictable future
- Because we cannot know both the position and velocity of an electron, we cannot predict the path it will follow
  - indeterminacy = indefinite future, can only predict probability
- The best we can do is to describe the probability an electron will be found in a particular region using statistical functions

Electron Energy

- Electron energy and position are complimentary
- For an electron with a given energy, the best we can do is describe a region in the atom of high probability of finding it – called an **orbital**
- Many of the properties of atoms are related to the energies of the electrons

Wave Functions

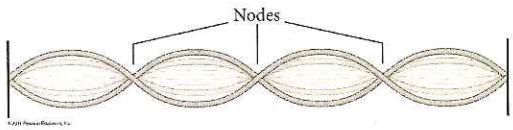
- Calculations show that the size, shape and orientation in space of an orbital are determined be three integer terms in the wave function
  - added to quantize the energy of the electron
- These integers are called **quantum numbers**
  - principal quantum number, n
  - angular momentum quantum number, l
  - magnetic quantum number, m_l

Principal Quantum Number, *n*

- Characterizes the energy of the electron in a particular orbital
- n can be any integer ≥ 1
- The larger the value of n, the more energy the orbital has
- The larger the value of n, the larger the orbital → Bigger atoms
- As n gets larger, the amount of energy between orbitals gets smaller

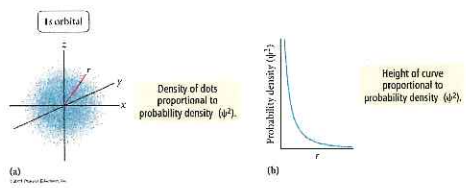
Probability & Radial Distribution Functions

- ψ^2 is the probability density
  - the probability of finding an electron at a particular point in space
- **Nodes** in the functions are where the probability drops to 0



Node like in the nucleus or around the nucleus since e^- cannot exsist in the nucleus

Probability Density Function



$n = 1, 2, 3 \ldots \infty$

$l = 0, 1, 2 \ldots n-1$

$m_l = \pm 0, \pm 1, \pm 2 \ldots \pm l$

$m_s = +\frac{1}{2}$ or $-\frac{1}{2}$

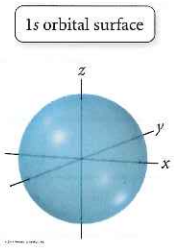
The Shapes of Atomic Orbitals

- The *l* quantum number primarily determines the shape of the orbital
- *l* can have integer values from 0 to $(n - 1)$
- Each value of *l* is called by a particular letter that designates the shape of the orbital
  - ***s*** orbitals are spherical
  - ***p*** orbitals are like two balloons tied at the knots
  - ***d*** orbitals are mainly like 4 balloons tied at the knot
  - ***f*** orbitals are mainly like 8 balloons tied at the knot

Know how to draw

l = 0, the **s** orbital

- Each principal energy state has 1 **s** orbital
- Lowest energy orbital in a principal energy state
- Spherical
- Number of nodes = (n – 1)



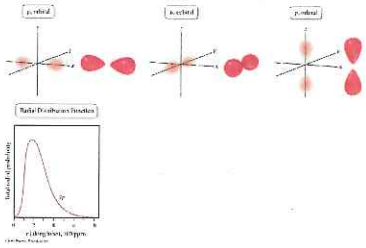
no nodal points

l = 1, **p** orbitals

- Each principal energy state above n = 1 has 3 **p** orbitals
  - m_l = -1, 0, +1
- Each of the 3 orbitals point along a different axis
  - p_x, p_y, p_z
- Two-lobed
- node at the nucleus, total of n nodes

Know how to draw

p orbitals

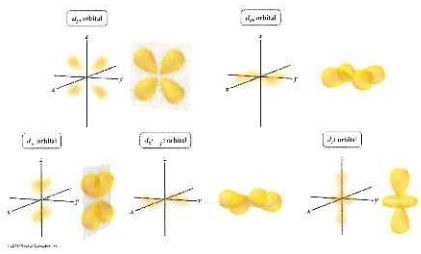


node at nucleus

l = 2, *d* orbitals

- Each principal energy state above n = 2 has 5 ***d*** orbitals
  - m_l = -2, -1, 0, +1, +2
- 4 of the 5 orbitals are aligned in a different plane
  - the fifth is aligned with the *z* axis, $d_{z\ squared}$
  - d_{xy}, d_{yz}, d_{xz}, $d_{x\ squared - y\ squared}$
- Mainly 4-lobed
  - one is two-lobed with a toroid
- Planar nodes
  - higher principal levels also have spherical nodes

d orbitals

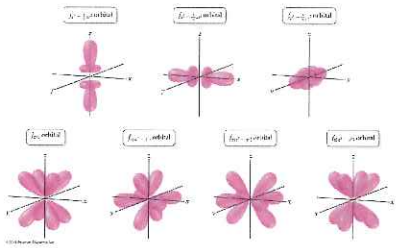


has a nodal plane

l = 3, *f* orbitals

- Each principal energy state above n = 3 has 7 ***d*** orbitals
  - m_l = -3, -2, -1, 0, +1, +2, +3
- Mainly 8-lobed
  - some 2-lobed with a toroid
- Planar nodes
  - higher principal levels also have spherical nodes

f orbitals



The End