HANDOUT SET

GENERAL CHEMISTRY I

Periodic Table of the Elements

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
	IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIIIB			IB	IIB	IIIA	IVA	VA	VIA	VIIA	VIIIA
1	1 **H** 1.00794																	2 **He** 4.00262
2	3 **Li** 6.941	4 **Be** 9.0122											5 **B** 10.811	6 **C** 12.011	7 **N** 14.0067	8 **O** 15.9994	9 **F** 18.9984	10 **Ne** 20.179
3	11 **Na** 22.9898	12 **Mg** 24.305											13 **Al** 26.98154	14 **Si** 28.0855	15 **P** 30.97376	16 **S** 32.066	17 **Cl** 35.453	18 **Ar** 39.948
4	19 **K** 39.0983	20 **Ca** 40.078	21 **Sc** 44.9559	22 **Ti** 47.88	23 **V** 50.9415	24 **Cr** 51.9961	25 **Mn** 54.9380	26 **Fe** 55.847	27 **Co** 58.9332	28 **Ni** 58.69	29 **Cu** 63.546	30 **Zn** 65.39	31 **Ga** 69.723	32 **Ge** 72.59	33 **As** 74.9216	34 **Se** 78.96	35 **Br** 79.904	36 **Kr** 83.80
5	37 **Rb** 85.4678	38 **Sr** 87.62	39 **Y** 88.9059	40 **Zr** 91.224	41 **Nb** 92.9064	42 **Mo** 95.94	43 **Tc** (98)	44 **Ru** 101.07	45 **Rh** 102.9055	46 **Pd** 106.42	47 **Ag** 107.8682	48 **Cd** 112.41	49 **In** 114.82	50 **Sn** 118.710	51 **Sb** 121.75	52 **Te** 127.60	53 **I** 126.9045	54 **Xe** 131.29
6	55 **Cs** 132.9054	56 **Ba** 137.34	57 **La*** 138.91	72 **Hf** 178.49	73 **Ta** 180.9479	74 **W** 183.85	75 **Re** 186.207	76 **Os** 190.2	77 **Ir** 192.22	78 **Pt** 195.08	79 **Au** 196.9665	80 **Hg** 200.59	81 **Tl** 204.383	82 **Pb** 207.2	83 **Bi** 208.9804	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
7	87 **Fr** (223)	88 **Ra** 226.0254	89 **Ac**** 227.0278	104 **Rf** (261)	105 **Db** (262)	106 **Sg** (263)	107 **Bh** (264)	108 **Hs** (265)	109 **Mt** (266)	110 (270)	111 (272)	112 ******* (277)						

*Lanthanides	58 **Ce** 140.12	59 **Pr** 140.9077	60 **Nd** 144.24	61 **Pm** (145)	62 **Sm** 150.36	63 **Eu** 151.96	64 **Gd** 157.25	65 **Tb** 158.925	66 **Dy** 162.50	67 **Ho** 164.930	68 **Er** 167.26	69 **Tm** 168.9342	70 **Yb** 173.04	71 **Lu** 174.967
**Actinides	90 **Th** 232.038	91 **Pa** 231.0659	92 **U** 238.0289	93 **Np** 237.0482	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (260)

Mass numbers in parenthesis are the mass numbers of the most stable isotopes. As of 1997 elements 110-112 have not been named.

***Peter Armbruster and Sigurd Hofman synthesized a single atom at the Heavy-Ion Research Center in Darmstadt, Germany in 1996. The atom survived for 280 μs after which it decayed to element 110 by loss of an α-particle

Chapter 1

Matter
Its Properties and Measurement

MATTER – ITS PROPERTIES AND MEASUREMENT

CHAPTER 1

INTRODUCTION

Lecture introduced the role of chemistry in the physical sciences and the *Scientific Method*. Further, we discussed a brief historical perspective of the origin of modern chemistry and tried to develop the idea that some of the theoretical principles we use today have evolved over hundreds or thousands of years. This chapter introduces many of the fundamental definitions used by scientists and emphasizes that the vocabulary is rigorously defined in the scientific context. Chapter 1 concludes with an in-depth discussion of the mathematical techniques used in the physical sciences.

GOALS

1. As a total part of your science education, you should understand how chemistry relates to the other physical and life sciences.
2. The historical perspectives of science is important in the total understanding of how and why science and technology is in its current state of development.
3. You should understand the general steps of the *Scientific Method*.
4. You must be comfortable with the terms and definitions outlined in lab, lecture, and text.
5. You must also be comfortable with measurements, the SI units of measure, and the SI base-units and prefixes.
6. The method of problem-solving known as *Dimensional Analysis*, *Factor-Label*, and *Conversion-Factor* is an extremely powerful technique that you should be familiar with. It is not absolutely essential that you master or even use the technique, but its superior strategies to working quantitative problems generally far outweigh all other techniques. Whenever possible you should avoid the *Ratio* method. On occasion, it is necessary to memorize an equation (*e.g.* temperature conversions) but this situation is rare if dimensional analysis is used.
7. While discussed mostly in the laboratory, knowledge of significant figures and how they are handled in calculations is essential and will be graded rigorously.

DEFINITIONS

You should have a working knowledge of at least these terms and any others used in lecture.

Scientific Method
Experiment
Hypothesis
Law
Theory
Property
Intensive
Extensive
Chemical property
Physical property
Chemical change
Physical change
Chemical reaction

Matter
Mass
Weight
Substance
Mixture
Homogeneous
Heterogeneous
Element
Atom
Compound
Molecule
Solid, liquid, gas
SI units

SI prefixes
Derived units
Dimensional Analysis
Conversion factor
Density
Percent
Significant figure
Most significant digit
Least significant digit
Precision
Accuracy
Systematic error
Random error

Some Definitions

Matter	Anything that occupies space
Mass	Quantity of matter
Weight	Effect of gravity on mass (not technically interchangeable with mass)
Substance	Matter with definite composition and distinct properties
Mixture	Combination of two or more substances in which each substance retains its chemical identity.

Heterogeneous	Individual substances retain their discreet boundaries
Homogeneous	Individual substances lose their discreet boundaries and are indistinguishable from the whole

The composition of a mixture is variable and the components can be separated by purely physical means.

Element	Pure substance that cannot be broken down into simpler substances by chemical means
Compound	Pure substance composed of 2 or more elements chemically united in fixed proportions by mass
Molecule	Single unit of a compound that retains the identity and properties of the compound
Atom	Single unit of an element that retains the identity and properties of the element

SI Base Units

Base Quantity	Unit	Symbol
Length	Meter	m
Mass	Kilogram	kg
Time	Second	s
Temperature	Kelvin	K
Quantity of Substance	Mole	mol
Electrical Charge	Coulomb	C
Electrical Current	Ampere	A
Luminous intensity	Candela	cd

SI Prefixes

Prefix	Symbol	Value	Example
Tera-	T	10^{12} (1,000,000,000,000)	1 terameter = 1 x 10^{12} m
Giga-	G	10^{9} (1,000,000,000)	1.21 gigawatts = 1.21 x 10^{9} Watts
Mega-	M	10^{6} (1,000,000)	1 megajoule = 1 x 10^{6} J
Kilo-	k	10^{3} (1,000)	1 kilometer = 1 x 10^{3} m
Deca-	D (not SI)	10	1 decagram = 10 g
Deci-	d (not SI)	10^{-1} (0.1)	1 decimeter = 0.1 m
Centi-	c (not SI)	10^{-2} (0.01)	1 centimeter = 0.01 m
Milli-	m	10^{-3} (0.001)	1 milligram = 10^{-3} g
Micro-	µ	10^{-6} (0.000001)	1 microliter = 10^{-6} L
Nano-	n	10^{-9} (0.000000001)	1 nanometer = 10^{-9} m
Pico-	p	10^{-12} (0.000000000001)	1 picomole = 10^{-12} mol
Femto-	f	10^{-15} (0.000000000000001)	1 femtogram = 10^{-15} g

Magnitude and Scientific (Exponential) Notation

1. Convert to scientific notation:

   100,000, precise to ±1

   ten thousand, precise to ± 1000

   0.000400

   0.0003

   275.3

2. Convert to exponential notation:

   175,906

   0.0000605

   two and a half million, precise to ±100

   two and a half billion, precise to ± one million.

3. Express each of the following in SI base units using scientific notation:

   432 kg

   624 ps

   1024 ng

   93,000 km, precise to ±10

   1 day

   0.0426 in.

SI Unit Conversions

1. Express each of the following in SI base units using scientific notation:

   a. 1 week

   b. 1.35 mm

   c. 15 miles

   d. 4.567 μs

   e. 6.45 mL

   f. 47 kg

2. The mass unit most commonly used for precious stones is the carat: 1 carat = 3.168 grains, and 1 gram = 15.4 grains. Find the total mass in kilograms (kg) of a ring that contains a 0.50 carat diamond and 7.00 grams of gold.

3. What is the total mass in grams, expressed in scientific notation with the correct number of significant figures, of a solution containing 2.000 kg of water, 6.5 g of sodium chloride, and 47.546 g of sugar?

Dimensional Analysis
A Chemical Problem

Carbon atoms weigh 12.01 g/mol. A mol is 6.022×10^{23} atoms. How many atoms of carbon are present in a 0.5 karat diamond? (A karat is 200 mg exactly.)

Dimensional Analysis
Example Density Problem

The density of lead is 11.34 g/mL. What Is the density of lead in pounds per cubic foot? Could you easily lift a 1.25 cubic foot block of lead? (Hint: calculate the weight of 1.25 cubic feet of Pb.)

Dimensional Analysis
A More Complex Problem

For a touch of reality (or absurdity) with particular appeal for the "trivia buff', try your hand at the following problem involving *dimensional analysis.* The problem is based on data given in the *Guinness Book of World Records.*

Problem:

The Amazon River has the greatest flow of any river in the world, discharging an average of 4,200,000 (2 significant figures) cubic feet of water per second into the Atlantic Ocean. If there are 1.48 x 10^{18} tons of water on Earth, how many years are required for the Amazon's flow to equal the Earth's water supply? (Assume a density of 1.0 g/cm^3 for all water.)

Mass-Mass Relationships in Calculations: Percentage and ppm

1. An aqueous solution of acetic acid ($HC_2H_3O_2$) is reported as 4.70% by mass (4.70% $^{w}/_{w}$). The density of the solution was determined to be 1.006 g/mL.

   a. What mass of the solution will contain 10.0 g of acetic acid?

   b. What volume (in microliters) of the acetic acid solution will contain 5.0 mg of acetic acid?

2. Many municipalities fluoridate their domestic drinking water prior to delivering the water to the customer. The fluoride helps to prevent dental caries (cavities) for people who live in areas where fluoride levels are naturally low. A typical concentration of fluoride is 0.70 ppm in the drinking water.

   At the typical concentration, what mass of fluorine will be found in a 12 oz (355 mL) glass of water. Assume the density of tap water is the same as pure water.

More Unit Conversions

1. What is the mass of 1 quart of water (1 L = 1.057 quarts)?

2. What is the mass of 1 quart of mercury (1 L = 1.057 quarts)?

3. Convert each of the following into SI units:

   a. engine displacement of 454 cubic inches

   b. car speed of 35 mph

   c. height of 6 feet 9 inches

   d. boulder mass of 227 pounds.

   e. gold nugget mass of 1.5 ounces

   f. light speed of $6.71 \text{ x } 10^8$ mph

   g. hike length of 11 miles

   h. car mileage of 32 miles/gallon.

Complex Units

1. A plastic block measures 15.5 cm by 4.6 cm by 1.75 cm, and its mass is 98.456 g. Compute the density of the plastic.

2. A penny has a diameter of 1.8 cm and a thickness of 0.15 cm, and its mass is 2.50 g. Compute the density of the penny (cylinder volume, $V = \pi r^2 h$).

3. Calculate the volume of an aluminum spoon whose mass is 15.4 g.

4. Calculate the volume of a quartz crystal of mass 0.246 g. (You may need to look up the density of quartz.)

Important Periodic Table Terms and Features

- **Regions**
  - Representative elements
  - Transition elements
  - Lanthanides
  - Actinides
- **Metals**
  - Location on Table
  - Properties
- **Nonmetals**
  - Location on Table
  - Properties
- **Metalloids (semimetals)**
  - Location on Table
  - Properties

- **Families (groups)**
  - Alkali metals
  - Alkaline earth metals
  - Chalcogens
  - Halogens
  - Noble gases
  - Coinage Metals
- **Periods**
  - Short periods
  - Long periods
- **Blocks**
  - *s*-block
  - *p*-block
  - *d*-block
  - *f*-block

Elements which naturally occur as homonuclear diatomic or polyatomic molecules

H_2 N_2 O_2 X_2 (X is all halogens) P_4 (one of several forms) S_8 (one of several forms)

C has several allotropes: graphite, diamond, fullerenes (e.g. C_{60} buckminsterfullerene or ‘bucky-ball”)

ORIGINS OF THE ELEMENT'S NAMES

Element	Symbol	Atomic Number	Origin of Name	Some Important Uses
Actinium	Ac	89	Greek, *Aktinos* (ray) for its radioactivity	Neutron source in nuclear energy applications
Aluminum	Al	13	Latin, *Alumen* (alum)	Lightweight alloys, wire, cans and foils, rugged reflective coatings on mirrors, abrasives, lasers (in the form of ruby and sapphire)
Americium	Am	95	Named after the American continent	Smoke detectors
Antimony	Sb	51	Latin, *stibium* = mark; Greek, *anthemonium*	Originally used as the sulfide by Asian women to darken their eyebrows. Name is possibly derived from the Arabic *al ithmid*, the name of Sb_2S_3 used for mascara. Used in semiconductor technology, metal alloys to increase hardness and strength, batteries, ordinance, flame retardants, paints, pharmaceuticals
Argon	Ar	18	Greek, *Aergon* (lazy one, no work, no action)	Used in incorrectly-named "neon" lights to produce red emission, electric light bulbs as an inert gas filler, fluorescent light bulbs, inert gas in inert-gas welding, semiconductor technology, inert gas atmosphere in chemical synthesis
Arsenic	As	33	Greek, *Arsenikos* (brave, male)	Semiconductors, bronzing, pyrotechnics, lasers
Astatine	At	85	Greek, *A-statos* (not lasting or not stable)	Radioactive element with 7.5 h half-life. There is probably less than 1 mg of the isolated element on Earth. There is no industrial use for astatine.
Barium	Ba	56	Greek, *Baryos* (heavy)	As the sulfate in x-ray analysis of the intestines, pyrotechnics, paint pigments
Berkelium	Bk	97	Named for Berkeley, the city of discovery	Currently, there are no industrial uses
Beryllium	Be	4	Greek, *Beryllos* for the mineral beryl	Hardener in metal alloys, ceramics, windows for x-ray sources
Bismuth	Bi	83	German, *Wiese* and *Muten* or more commonly accepted *Weisse Masse* meaning white mass	OTC medications such as PeptoBismol$^{\circledR}$, iron alloys, catalysts, thermocouples, fire extinguishers
Bohrium	Bh	107	Named in honor of Niels Bohr	No industrial use
Boron	B	5	Arabic, *Bauraq*; Persian, *Burah* for borax, $Na_2B_4O_7 \cdot 10H_2O$ where boron is found	Insecticides (as boric acid), pyrotechnics, borax, paint, ceramics, borosilicate glass (*e.g.*, Pyrex and Kimax), nuclear reactors,
Bromine	Br	35	Greek, *Bromos* (bad odor or stink)	Flame retardants, dyes, pharmaceuticals, photography (as AgBr), fumigants, pesticides, pool water disinfection
Cadmium	Cd	48	From *cadmium fornacum* (furnace calamine). Calamine ($ZnCO_3$) was a mineral found in *Kadmeia* in Ancient Greece.	Specialty solders and other alloys, iron coating to prevent rust, batteries, nuclear control rods, television CRT phosphors, paint pigment
Calcium	Ca	20	Greek, *kylix*; Latin, *calx* (chalk)	Cement, reducing agent, used in the preparation of various alloys, formerly used in vacuum tubes
Californium	Cf	98	Named for California, the state of discovery	^{252}Cf foils are used as a source of fission fragments for research purposes, neutron source for detectors and scintillation counters, moisture gauges for the determination of water and oil-bearing layers in oil wells

Element	Symbol	Atomic Number	Origin of Name	Some Important Uses
Carbon	C	6	From Latin, *carbo* (charcoal)	As graphite for lubrication, as a fossil fuel (coal), for hardening steel, radiological dating, basic “building block” element of millions of organic compounds
Cerium	Ce	58	Latin, *Ceres* the Roman goddess of harvest and the name of the first asteroid discovered	Oxide is one component of incandescent gas mantles (Coleman lanterns), catalyst in "self-cleaning" ovens, oxide used as a polishing compound, nuclear applications
Cesium	Cs	55	Latin, *Caesius* (blue of the upper part of the firmament) from the blue color of the vapor excited to incandescence	Catalyst, photoelectric effect detectors, atomic clocks, IR emitters, IR optics
Chlorine	Cl	17	Greek, *khloros* (yellow-green) for the color of the gas	Disinfection of water, textiles, pharmaceuticals, insecticides, herbicides, paints, plastics, solvents, bleach
Chromium	Cr	24	Greek, *khroma* (color) for the many colored compounds it produces	Metal alloys, hardening steel, corrosion resistant surfaces, gives glass a green color and rubies their red color, paint pigments, leather tanning
Cobalt	Co	27	German, *kobold* (evil sprite)	Paint pigments, strong Alnico magnets, high strength alloys for turbine engines, stainless steel, ^{60}Co is used for radiotherapy in cancer treatment
Copper	Cu	29	Greek, *Kyprion*; Latin, *Cuprum*	Wire, coins, the electrical industry, copper sulfate is used to kill water-borne fungus and algae
Curium	Cm	96	In honor of Marie and Pierre Curie	Alpha particle source, used aboard Mars landers in the Alpha Proton X-Ray Spectrometer
Darmstadtium	Ds	110	Named for Darmstadt, Germany, city of first synthesis	Since only a few atoms have ever been made, this element has no uses
Dubnium	Db	105	Named for Dubnia, the location of its first production	No known use
Dysprosium	Dy	66	Greek "*dysprositos*" (hard to obtain) since it was difficult to separate it from a holmium mineral	Possible use in nuclear reactor control, solid-state lasers
Einsteinium	Es	99	Named in honor of Albert Einstein	So little has been made that there are no uses
Erbium	Er	68	Named after the town Ytterby, Sweden since it was found in the mineral *ytterite*	Glass and ceramics, nuclear industry, reduces brittleness of hard metals
Europium	Eu	63	Named for the continent	Alloys, nuclear reactor control, with yttrium makes the phosphor that produces the red color in television CRTs, lasers
Fermium	Fm	100	Named in honor of Enrico Fermi	So little has been made that there are no uses
Fluorine	F	9	Latin, *Fluere* (flow) for the ease of melting fluorspar	In domestic drinking water to prevent tooth decay, plastics, Teflon, as hydrofluoric acid for decorative glass etching, refrigerants
Francium	Fr	87	Named for France, homeland of its discoverer	So little of the element exists that no uses are possible
Gadolinium	Gd	64	After Sir Johan Gadolin of Turku, the first to study the mineral gadolinite	Contrast dyes in MRI analysis, television cathode ray tubes, alloys, compact disks, superconductors

Element	Symbol	Atomic Number	Origin of Name	Some Important Uses
Gallium	Ga	31	Latin, *France* for its country of discovery	Semiconductors, solid-state lasers, alloys, neutrino detectors
Germanium	Ge	32	After Germany, its country of discovery	Transistors and other semiconductors, alloys, phosphor in some fluorescent lamps, catalysis, optics – especially far-IR, radiation detectors
Gold	Au	79	Latin, *Aurum* (yellow) and from Aurora, goddess of dawn	coinage and currency, jewelry, decoration, dentistry, IR mirrors, electroplating, photography, anti-arthritis drugs, electrical connectors
Hafnium	Hf	72	Latin for Copenhagen, the city of discovery	Nuclear control rods, gas filled and incandescent lamps
Hassium	Hs	108	Latin, *Hassias* (Germany). Named for the German State of Hesse	Since only a few atoms have ever been synthesized, there is no use for the element.
Helium	He	2	Greek, *Helios* (Sun) for where helium was discovered (N. Lockyer)	Balloons, lighter-than-air aircraft, inert gas shield for arc welding (especially Al), coolant gas for nuclear reactors, in liquefied form a cryogenic gas, mixed-air breathing gas modifier for ultra-deep divers, carrier gas in gas chromatography
Holmium	Ho	67	Named for *Holmium*, the ancient name of Stockholm, Sweden	Alloys, optics
Hydrogen	H	1	Greek, *Hydros* (water) and *Gen* (producing)	Hydrogenation of fats and oils, production of ammonia, rocket fuel for the space shuttle, reduction of metallic ores, formerly used in lighter-than-air craft, cryogenics, tritium (^{3}H) is used medicinally (*e.g.*, in PET scan technology) and in thermonuclear weapons.
Indium	In	49	Latin, *indicum*; Greek, *indicon* for the indigo line emission from the excited state atom.	Metal alloys, semiconductors, corrosion resistant mirrors
Iodine	I	53	Greek, *Ioeides* (violet) for the color of its vapor	Medicinal disinfectant, pharmaceuticals, photography (as AgI), nutritional supplement (as KI in table salt)
Iridium	Ir	77	Greek, *Iris* (rainbow) because its salts show a variety of colors	High temperature crucibles and apparatus, electrical and electronics applications, hardening agent for platinum, alloys with osmium for pen tips and compass bearings
Iron	Fe	26	Arabic(?), *Ferrum*; Anglo-Saxon, *Iren*	Probably smelted first, in extreme secret, by the Egyptians and/or Hittites in 3000 B.C. By 1200 B.C., the start of the Iron Age, smelting and use of iron extensively spread. Used in steel and thousands of other uses
Krypton	Kr	36	Greek, *Kryptos* (hidden) for the gas “hiding” in distilled liquefied air	Used in place of xenon in some photographic flash units, lasers, fill-gas in some fluorescent light bulbs
Lanthanum	La	57	Greek, *Lanthanos* (escapes notice) since it was hidden in the mineral cerite for nearly 40 years after a compound of cerium was discovered	Metal alloys, lighter flints
Lawrencium	Lr	103	Named in honor of Ernest Lawrence	So little has been made that there are no uses

Element	Symbol	Atomic Number	Origin of Name	Some Important Uses
Lead	Pb	82	Latin, *Plumbum*	Ancient uses include water-pipes, writing tablets, coins, and cooking utensils. Modern uses include: metal and the oxide in lead-acid (automobile) batteries, ammunition, radiation shielding around x-ray equipment and nuclear reactors, paints, the oxide is used in producing fine "leaded-crystal" glass, solder
Lithium	Li	3	Greek, *Lithos* (stone)	Some salts are used to treat manic-depressive disorders, lubricants, CO2 scrubber in manned space craft, aircraft alloys, organic chemistry synthesis, batteries
Lutetium	Lu	71	Latin, ancient name of Paris	Catalysts, alloys
Magnesium	Mg	12	From *magnesia*	Used in flares, pyrotechnics, military ordnance, alloys for use in aircraft, automobile engines, and wheels, OTC pharmaceuticals such as the hydroxide (milk of magnesia) and sulfate (Epsom salts)
Manganese	Mn	25	From *magnesia nigri* (black magnesia), which distinguishes it from *magnesia alba* (white magnesia), a source of magnesium	As the permanganate a strong oxidizing agent, in steel alloys to improve strength, wear resistance, and hardness, magnets, batteries, glasses and ceramics, responsible for the color of amethyst, required for biological utilization of vitamin B_1
Meitnerium	Mt	109	Named for early 20th century physicist Lise Meitner	No industrial use
Mendelevium	Md	101	Named for Dmitri Mendeleev	No industrial use
Mercury	Hg	80	Latin, *Hydrargyrum* from Greek, *Hydro-argyros* (water-silver)	Thermometers, barometers, diffusion pumps, and many other instruments, mercury switches and other electrical apparatus, batteries, mercury-vapor lamps and advertising signs, dental amalgams
Molybdenum	Mo	42	Greek, *Molybdos* (lead)	Hardens steel alloys, aircraft industry, essential for nutrition, lubricants
Neodymium	Nd	60	From the Greek *neos didymos* (new twin)	Discovered in 1885 by Austrian chemist Carl Auer von Welsbach who separated didymium into two elements, one of which he called neodymium (new twin). Used in powerful permanent magnets, colorant for glasses and glazes, metal alloys, welding goggles, solid-state lasers, calibration glasses in spectroscopy.
Neon	Ne	10	Greek, *Neos* (new)	Largest use in making neon emission tubes for signs, lasers, cryogenic refrigerant when liquefied
Neptunium	Np	93	From the planet Neptune – next element after Uranium and next planet after Uranus.	By product of nuclear fission, ^{237}Np is a component in neutron detection instruments
Nickel	Ni	28	From German *Kupfer-nickel* (devil copper) since the ore resembles copper ore.	stainless steel and other corrosion-resistant alloys, coinage, electroplating, hydrogenation catalysts, batteries (*e.g.*, Ni-Cd and nickel metal-hydride batteries)
Niobium	Nb	41	From Niobe, the daughter of Tantalus in Greek mythology	A component of some stainless steels, other alloys, welding rods, nuclear reactors, wire for superconducting magnets, tongue and navel studs
Nitrogen	N	7	Greek, *niter* = saltpeter	millions of chemical compounds, all plant and animal cells, fertilizers, ammonia

Element	Symbol	Atomic Number	Origin of Name	Some Important Uses
Nobelium	No	102	Named in honor of Alfred Nobel	So little has been made that there are no uses
Osmium	Os	76	Greek, *Osme* (odor) for the disagreeable odor of OsO_4	Almost entirely used to produce very hard alloys with other metals of the platinum group, for fountain pen tips, instrument pivots and bearings, and electrical contacts
Oxygen	O	8	Greek, *oxys* (acid) and *genes* (forming)	Respiration, combustion, millions of chemical compounds
Palladium	Pd	46	Greek, *Pallas* goddess of wisdom and name of second asteroid discovered	Jewelry (e.g., production of "white" gold), catalyst, dentistry, electrochemical electrode applications
Phosphorus	P	15	Greek for light producing	All plant and animal cells, fertilizers, detergents. Discovery of phosphorus is the earliest for which its discoverer is known.
Platinum	Pt	78	Spanish, *platina* (silver)	Jewelry, jet engines and missiles, corrosion resistance, catalysts
Plutonium	Pu	94	From the planet Pluto – next element after Neptunium and next planet after Neptune.	Nuclear power plants, nuclear weapons
Polonium	Po	84	Named after Poland, homeland of its discoverer (Marie Curie)	Neutron source in nuclear energy, "static-free" brushes, nuclear reactor fuel in some spacecraft
Potassium	K	19	English, *potash* German, *kalium*	Fertilizers, present in thousands of compounds, essential for proper neural activity
Praseodymium	Pr	59	Greek, *Praseios* (green color of a leek) for the color of most of its salts	Alloys, Misch metal in cigarette lighters, glasses and ceramics, carbon arc-lights
Promethium	Pm	61	From mythology, Prometheus, the god who stole fire from heaven and gave it to humans	Currently little industrial use. All promethium on Earth is synthetic.
Protactinium	Pa	91	Greek, *Protos* (prior) for being the parent element that produces the daughter element, actinium.	No known industrial use
Radium	Ra	88	Latin, *Radius* (ray) for the sulfate salt which glows in the dark (disc. by M. and P. Curie)	Formerly used for self-luminescent watch and clock faces, formerly used in cancer radio-therapy
Radon	Rn	86	Latin, *Radius* (ray) for being produced in air exposed to radium	Little industrial use, gas with highest known density
Rhenium	Re	75	Latin for the Rhenany-Reinland, birthplace of Ida Tacke, one of the discoverers	90% of rhenium is used for catalysts, tungsten and molybdenum-based alloys, filaments for mass spectrographs and ion gauges, superconductors, thermocouples
Rhodium	Rh	45	Greek, *Rhodon* (rose) for its rose-colored salts	Automobile catalytic converters, high temperature thermocouple elements, high temperature crucibles, electrical contact material, catalysts, jewelry
Rubidium	Rb	37	Latin, *Rubidus* (deepest red) from the strong red emission of excited gas-phase rubidium atoms	possible use in "ion engines" for space vehicles, photocells, specialty glasses

Element	Symbol	Atomic Number	Origin of Name	Some Important Uses
Ruthenium	Ru	44	Latin, *Russia*	Alloyed into platinum and palladium to harden the metal, added to titanium for improved corrosion resistance of titanium, catalyst
Rutherfordium	Rf	104	In honor of New Zealand physicist, Ernest Rutherford	No industrial use
Samarium	Sm	62	From Samarskite, the mineral source of samarium	High-field magnets, carbon-arc lighting, lasers, infrared detectors, catalysts
Scandium	Sc	21	For Scandinavia, homeland of Lars Fredrik Nilson (1840-1899), discoverer of the element	No significant industrial use
Seaborgium	Sg	106	After the American chemist, Glenn Seaborg	No industrial use
Selenium	Se	34	Greek, *Selene* (Moon)	Semiconductor technology, photocells, photography, stainless steel alloys
Silicon	Si	14	Latin, *silex* (flint)	Semiconductors, computer chips, abrasives, tools, water repellents
Silver	Ag	47	Latin, *argentum*	Coinage, photographic chemicals, jewelry silverware, electrical contacts, batteries
Sodium	Na	11	Hebrew, *Neter*; Latin, *Nitrium*; German, *natrium*	Ubiquitous, many compounds such as salt (NaCl), lye (NaOH), essential for proper neuronal activity
Strontium	Sr	38	From the mineral *strontianite* found in the town of Strontian, Argyleshire, Scotland	Specialty optical glasses, pyrotechnics to produce red fireworks, road and signal flares, nuclear reactors, television CRTs
Sulfur	S	16	Sanskrit, *sulvere*	Gunpowder, automobile tires, sulfuric acid, paper, fumigants, fertilizers
Tantalum	Ta	73	From mythology, *Tantalus*, son of Jupiter, was condemned to hell, standing to his neck in water. But the water sank when he stooped to drink. Ta_2O_5 does not take in water, nor dissolve in acids.	Optical glass with high refractive index, ductile steel alloys with high melting point and high strength, military ordnance, nuclear reactors, high biological tolerance makes it suitable for surgical use, electronic components
Technecium	Tc	43	Greek, *technitos* (artificial). Formerly masurium, Ma	^{95}Tc radioactive tracer, corrosion inhibition in steel, superconductive below 11 K, medical imaging agents
Tellurium	Te	52	Latin, *Tellus* (Earth)	Semiconductors, metal alloys, ceramic and glass colorants
Terbium	Tb	65	Named after the town Ytterby, Sweden since it was found in the mineral *ytterite*	Alloys, lasers, semiconductor technology
Thallium	Tl	81	Latin, *Thallus* (sprouting green twig) for the brilliant green emission from the excited thallium vapor	Extremely toxic metal. Used in IR detectors and visible light photocells, formerly used in rodenticides and pesticides, specialty glasses

Element	Symbol	Atomic Number	Origin of Name	Some Important Uses
Thorium	Th	90	After *Thor*, Norse god of war	Preparation of the "Welsbach mantle", used for portable gas lights (*e.g.*, Coleman lanterns), alloys, laboratory crucibles, high refractive index glasses, catalysts.
Thulium	Tm	69	Named after Thule, the ancient name of Scandinavia	Alloys, little industrial use
Tin	Sn	50	Latin, *stannum*	Alloys (pewter, bronze, solder), coating of steel for tin cans
Titanium	Ti	22	From the *Titans* (giants) of Greek mythology	Aerospace alloys, medicinal alloys (e.g., artificial hip joints), paint pigments (the largest use of TiO2), artificial gemstones, ship hardware continually exposed to sea water
Tungsten	W	74	Swedish *tung sten* (heavy) German, *wolfram*	Highest melting point element, light bulbs, dental drills, steel alloys
Uranium	U	92	Greek, *Uranos* god of heaven. Named after the planet Uranus	Fuel in nuclear fired power plants, nuclear weapons, starting material for production of Pu
Vanadium	V	23	Named for *Vanadis* (or Freya), the Norse goddess of beauty because of the many colored compounds vanadium forms.	Spring and stainless steels, nuclear industry, ceramics, catalysts, superconductive wires
Xenon	Xe	54	Greek, *Xenos* (stranger)	Photographic flashes and strobes, lasers, UV sterilizing lamps, ion-engines for spacecraft
Ytterbium	Yb	70	Named after the town Ytterby, Sweden since it was found in the mineral *ytterite*	Lasers, steel alloys
Yttrium	Y	39	From the mineral *ytterite*, discovered by Johan Gadolin (also gadolinium)	Phosphors containing yttrium produce the red color in television CRTs, high temperature superconductors, simulated diamonds, lasers, catalyst, alloy to strengthen some soft metals
Zinc	Zn	30	German, *zinke* (spike); Persian, *seng* (stone)	Galvanized nails (zinc-coated), die castings, pigments, cosmetics, some proteins
Zirconium	Zr	40	Arabic, *Zerk* (precious stone)	Used in crucibles for corrosive chemicals, cladding for nuclear reactor fuel rods, steel alloys, superconducting wire (with niobium), gems (as the oxide)

References

__, *Handbook of Chemistry and Physics*, Chemical Rubber Co., 1985

Bowden, Mary Ellen *Chemical Achievers*, Chemical Heritage Foundation, 1997

Curie, Pierre; Curie, Marie S. "On a New Radioactive Substance Contained in Pitchblende" *Comptes Rendus 127*, **1898**, 175-8. 1898. (translated/reprinted in Boorse, H.A. and Motz, L., eds., *The World of the Atom*, **1** Basic Books, NY, 1966)

Döbereiner, J.W. "An Attempt to Group Elementary Substances according to Their Analogies" *Annalen der Physik und Chemie 15*, **1829**, 301-7

Holden, N.E. http://www.nndc.bnl.gov/content/elements.html (National Nuclear Data Center, Brookhaven National Laboratory)

Ihde, Aaron J. *The Development of Modern Chemistry*, Dover, 1984

James, L.K. *Nobel Laureates in Chemistry 1901 - 1992* American Chemical Society & Chemical Heritage Foundation,1993

Lecoq de Boisbaudran, P.E. "About a New Metal, Gallium" *Annales de Chimie 10*, **1877**, 100-141

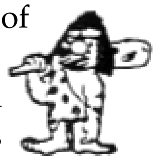
Weeks, Mary Elvira "Discovery of the Elements", *Journal of Chemical Education*, 1945

Winter, Mark www.webelements.com (University of Sheffield, England)

A Brief History of Chemistry

The Prehistoric Period to A.D. 500

Since one of the objectives of chemistry is to acquire knowledge of the properties of matter, every bit of information that prehistoric man found out about the materials around him was, in a sense, a contribution to chemical knowledge. When man first learned, for example, that wood burns and stone does not, and then passed this information along to his fellows, he took the first step along the road toward chemical knowledge.



During these early centuries of man's life on earth., he acquired a great deal. of useful information about the properties of matter. We find direct and indirect evidence in early Biblical writings and other records that, as the years passed, man continued to extend his knowledge of matter and to put it to more and more uses. He learned to dig metals out of the earth, particularly gold, silver, and copper, and to fashion them into useful articles. He discovered how to make glass long before he fully understood its chemical make-up and behavior. Many other substances, such as medicines, oils, and dyes from plants, came into widespread use during this period. Apparently, however, there was no successful attempt to classify or correlate this new-found knowledge, and little progress was made toward a science of chemistry.

The Period of Alchemy, A.D. 500-1600

During these years, many capable and inquiring men turned their attention to the direct observation of matter. The nonsystematized studies by these medieval investigators is called *alchemy*.

Some alchemists made an effort to discover a method of changing base metals, such as lead, into noble metals, such as gold. They engaged in a great deal of experimentation, but they destroyed much of the potential value of their findings by keeping them secret or by misrepresenting the results they actually obtained. The alchemists succeeded, however, in preparing many new substances and in developing scores of useful pieces of apparatus and experimental skills.

The Medical-Chemical Period, 1100-1750

In some ways this period was similar to the period of alchemy. Men experimented widely and began to do some theorizing. In their search for effective medicines, they prepared and purified many new chemical substances, and therefore made valuable contributions to the future. It was during these years that the experimental and the theoretical approaches were successfully wedded for the first time. Francis Bacon first urged that the experimental study of nature be combined with theoretical interpretation. Then, as other scholars adopted Bacon's approach, scientific progress began its swift acceleration toward the spectacular growth rate it has achieved in this century.

Period of Phlogiston Theory, 1700-1777

The scientists of the eighteenth century made an intensive study of the process of burning and developed several theories to explain it. In *1702* Georg Ernst Stahl, a German chemist, proposed that some substance actually was released during the burning of combustible matter. This substance was called 'phlogiston," from the Greek word *phlogistos*, which translates as *flammable*. Stahl's theory was strongly defended for 75 years. One reason for its wide acceptance was the failure of chemists to use weighing devices to determine the exact weights of materials before and after they were burned. This is an interesting example of a theory that seems correct and consistent with a large body of observed facts, but is proved incorrect by more careful quantitative experimental work.

The Modern Period, l777-Present

The modern period of chemistry dates from the work of the French chemist Antoine Lavoisier (1743-1794). We shall be hearing a great deal more of Lavoisier in this book, so let us pause here a moment to give a brief account of this life. He was born in Paris, and at the age of 23 was awarded a gold medal by the Academy of Sciences (of Paris) in recognition of a report he compiled on the problem of city lighting. Most of his life was spent in Paris, and most of his research was done at the Sorbonne. His work on combustion we shall discuss later. In addition, he developed a theory of acids which, though erroneous, was a significant advance over earlier ideas. He also drew up a nomenclature of the elements. Apart from his scientific researches, he served as advisor on many public committees both national and municipal, activities, however, which during the French Revolution led to his death by the guillotine in 1794.

Lavoisier made the first extensive use of weighing devices (balances) in chemical studies, which led to the discovery of the fundamental importance of the conservation of the mass of matter in chemical reactions. Now chemists could measure accurately the amount of matter used and produced in chemical reactions. And on the basis of these exact experimental data, they could proceed to develop acceptable theories and precise chemical laws. This great step forward raised chemistry to the status of an exact science. During the succeeding years, development sped along at an ever-increasing rate, providing the impetus for the discoveries that mark our own highly advanced technical age.

Chapter 2

Atomic Theory

ATOMIC THEORY

CHAPTER 2

INTRODUCTION

The history of the modern atomic theory spans nearly 2000 years. However, it is only in the last 200 years that there has been a good understanding of the atom and even still new discoveries about microscopic atomic structure are being made. It is important as a scientist to know the people involved in discovery and understand the experiments performed which led to our current understanding of the atom.

With the discovery of the subatomic particles such as the electron, proton, and neutron we can explain isotopes, bonding, and many other atomic properties. Further, knowledge of isotopes helps us to understand why atoms do not have integral masses - the measured atomic mass of an element is a weighted-average atomic mass of the isotopes. Finally, we can use Avogadro's number to relate atomic masses on the periodic table to molar masses of each element. Through Avogadro's number, we can count exceeding large numbers of atoms using a more convenient unit, the mole.

GOALS

1. You should know the historical perspectives of atomic theory, with particular attention to those scientists who made the most significant contributions.
2. Several *laws* have been introduced of which you should have a working knowledge.
3. Many features of the periodic table have been introduced including many elemental symbols and names. You should have an active working familiarity with any element discussed in lecture or lab. Remember that only symbol, name, and rough positioning in the periodic table are important. Do not memorize atomic number or mass.
4. You should understand what isotopes are and how they are symbolically represented. You should also know how they were first discovered and what their effect is on the calculated atomic mass.
5. Possibly the most important concept presented in this chapter is that of the *mole concept*. Remember that a mole is just like a dozen. Whereas a dozen cookies is 12 cookies, a mole of atoms is 6.022 x 10^{23} atoms. You must be able to work with numbers of atoms, moles of atoms, and masses of atoms routinely and with little difficulty.

DEFINITIONS

You should have a working knowledge of at least these terms and any others used in lecture.

Atomic Theory
Anode
Cathode
Cathode Ray Tube
Cathode rays
Canal rays
Anode rays
x-rays
Electrons
Protons
Neutrons
Radiation
Radioactivity
α-particle
β-particle
γ-ray
"Plum-pudding" model
"Nuclear" model
"Oil-drop" experiment
"Gold-foil" experiment
Isotope
Nuclide
Abundance
Atomic mass unit (amu, u)
Mass number
Atomic number
Avogadro's number
Mole
Molar mass
Law of Conservation of mass
Law of Definite Proportions

Who Did It?

Date	Person	Claim(s) to Fame
~400 B.C.		Democritus *Atomos* theory and *Atomism*
1600's	Isaac Newton	Separated sunlight into the rainbow (continuum) with a prism.
1799	Joseph Proust	Contributed to the Law of the Constant Composition (aka Definite Proportions)
Early 1800's	Humphrey Davy	Hypothesized that substances are held together by electrical forces.
1808	John Dalton	*Atomic Theory*. First experimentally verifiable theory of the atom.
1832	Michael Faraday	(student of Davy) Determined the relationship between amount of charge used by a chemical reaction (coulombs) and extent of reaction (moles). (96500 ^{C}/mol)
1859	J. Plücker	Built the first *Cathode Ray Tube*.
1865	James Maxwell	Described light as *Electromagnetic Radiation*.
1895	Wilhelm Röntgen	Discovered x-rays by striking metal targets with *cathode rays* (see also J.J. Thomson, 1897). One of today's units of radioactivity is named after Röntgen.
1896	Eugen Goldstein	Observed that in a *cathode ray tube* (CRT), determined that positive "rays" were generated from residual gases in the CRT. The "canal rays" had different *e/m* ratios depending on gas. Greatest *e/m* ratio when hydrogen residual gas - hypothesized a fundamental particle, the *proton*.
1897	Joseph John Thomson	Using the *cathode ray tube* (CRT), discovered the *electron* (cathode rays), a new subatomic particle. Determined *e/m* = -1.76 x 10^{8} ^{C}/g
1900	Joseph John Thomson	"Plum-pudding" model of the atom.
1900	Max Planck	Suggested wave-particle duality of EM radiation - light emitted in packets of energy called *quanta* (singular *quantum*). Energy of quantum: $E = h\nu$
1903	Antoine Henri Becquerel	Discovered that uranium emitted high energy radiation.
1905	Marie and Pierre Curie	Discovered and characterized many new *radioactive* substances. Marie and Becquerel shared the 1905 Nobel Prize in physics.

1905	Albert Einstein	Explanation of the *photoelectric effect* using packets of quanta called *photons*.
1908	Robert Millikan	"Oil-drop experiment". Determined the charge on the electron. Also determined the accepted value of Avogadro's number.
1910	Ernest Rutherford	(student of Thomson) Introduced the "nuclear model" of the atom using data from the *gold-foil* experiment. (with Ernest Marsden and Ernest Geiger)
1913	Niels Bohr	Postulated that electrons orbit around nucleus like planets around the sun. Further postulated that the orbits were *quantized*. Intriguing model since the equations gave exactly the right answer for hydrogen atom and was very close for He^+ and adequately explained the origin of line spectra for gases. Unfortunately, the model could not be extended to "many-electron" atoms.
1922	Otto Stern W. Gerlach	Showed that electrons possess two and only two spins - observation known as the *Stern-Gerlach effect*.
1923	Louis de Broglie	Explained why the orbits of the electrons in the hydrogen atom are quantized using the reasoning of Max Planck...if photons can behave like particles or waves, why not electrons also? While Bohr's model is not correct, de Broglie's equation explains other behaviors observed for moving particles.
1932	James Chadwick	Bombarded different metal targets with α-particles. Called the heavy neutral particles emitted from target *neutrons*.
1926	Erwin Schrödinger	Developed a new branch of mathematics and physics and presented the *Quantum Mechanical* model of the atom. Electrons are around the nucleus in regions of high *probability* called *orbitals*. Electron energy levels quantized. Quantum mechanics not only accurately models the line spectrum observed for hydrogen, but also explains effects observed in many-electron atoms.
1927	Werner Heisenberg	Asserted that it is not possible to know the position and velocity of an electron simultaneously. By virtue of measuring one the other changes.

Dalton's Atomic Theory

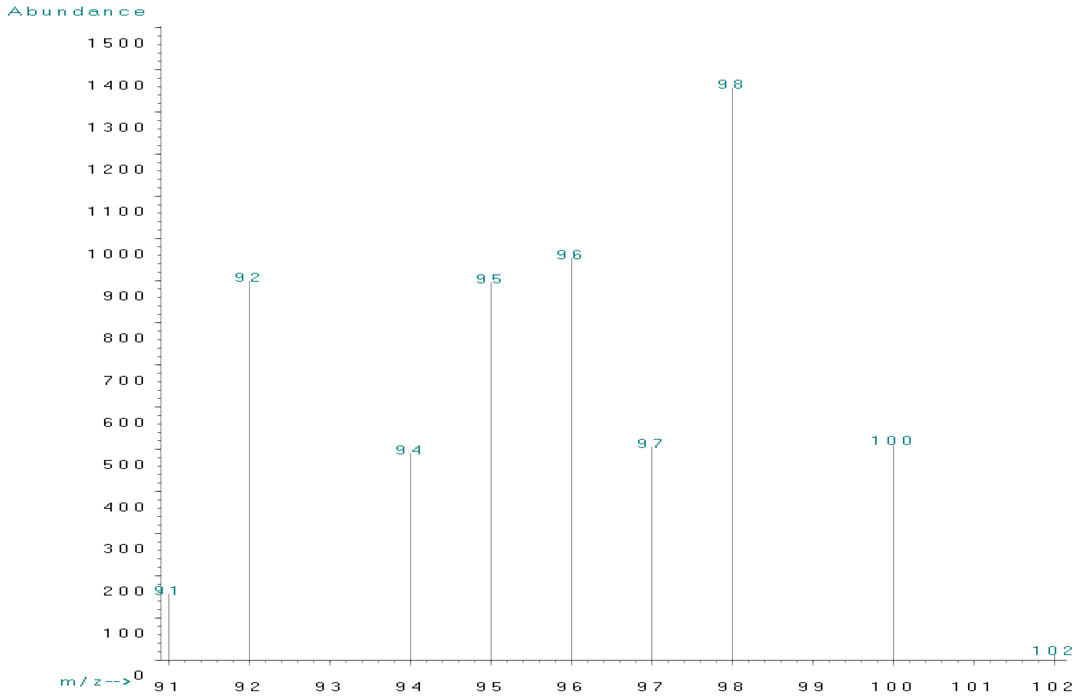
1. **An element is composed of microscopic indivisible spheres.**
   Good postulate for the time. With the discovery of the electron (ca 1906) this point must be amended.

2. **All atoms of the same element are identical**
   Again, a good postulate for the time. It wasn't until 1932 that neutrons were discovered verifying the existence of nuclear isotopes. All atoms of the same element are the same or differ only by their number of neutrons.

3. **Atoms combine in simple whole number ratios.**
   The Law of Definite Composition (or Proportions). For the most part this axiom has remained unchanged since the time of Dalton.

4. **The numbers and kinds of elements in a given compound are constant.**

5. **Chemical reactions are the combination and rearrangement of atoms but not their creation nor destruction.**
   This is the Law of Conservation of Mass. The only modification to the Law of Conservation of Mass is the qualification that the reaction be a chemical reaction. In nuclear reactions energy but not mass is conserved (according to Einstein mass and energy are related entities).



The mass spectrum for molybdenum metal is shown above as acquired on Pepperdine's mass spectrometer. The x-axis of the mass spectrum is actually the mass-to-charge ratio for each isotope; however, in this case *m/z* is the same as atomic mass units. The y-axis is named *Abundance* but actually refers only to the relative number of each isotope in the mass analyzer. Recall from lecture discussion and reading that the abundance of an isotope is proportional to the intensity of the mass line in the mass spectrum.

Use a ruler and the exact mass information in the table below to determine the percentage abundance of each significant isotope of molybdenum. Based on the masses and abundances, calculate the weighted-average atomic mass of the element Mo. Explain any differences in the calculated atomic mass from the atomic mass listed on the periodic table.

Isotope	Isotopic Mass (u)	Percentage Abundance
92	91.9063	
94	93.9047	
95	94.90584	
96	95.9046	
97	96.9058	
98	97.9055	
100	99.9076	

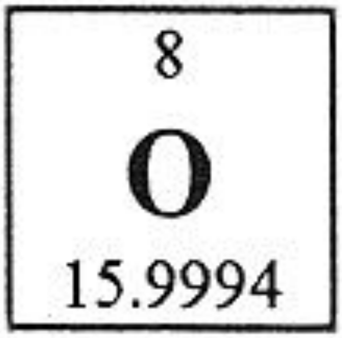
Average Mass of an Element

The element neon is composed of 3 isotopes. The abundances and isotopic masses are listed in the table below. What is the weighted-average atomic mass of neon in atomic mass units, u?

atomic mass (u)	abundance (%)
19.992	90.92
20.994	0.257
21.991	8.82

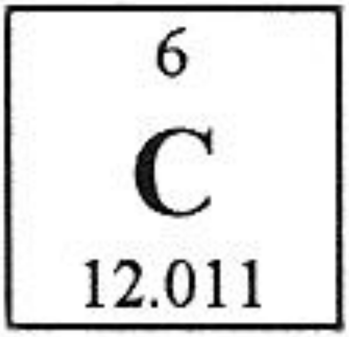
Brief and Mostly Accurate History of the AMU Scale

IN THE VERY EARLY 18TH CENTURY, chemists (at least the equivalent of the day) decided that since hydrogen was the lightest of all elements (or the least dense, if you prefer), then the unit of atomic mass should be the mass of the hydrogen atom. So from then on, all atomic masses were measured as multiples of the hydrogen atom. Since 1.0 g of hydrogen combined with 35.5 g of chlorine to make hydrogen chloride HCl, the atomic mass of chlorine was taken to be 35.5 amu.



Then a complication: at first, chemists thought that the formula of water was HO. Since 1.0 g of hydrogen combined with 8.0 g of oxygen, it was thought for some time that the atomic weight of oxygen was 8.0 amu. Later, it became clear that the formula of water is H_2O. So if one oxygen combined with two hydrogens, then 2.0 g of hydrogen combined with 16.0 g of oxygen; thus the mass of an oxygen atom must be $2 \times 8.0 = 16.0$ amu. The rest of the atomic masses were derived in the same way, from a knowledge of their combining-weights, coupled with a knowledge of the formula of the relevant compounds.

Thus the – archaic – interpretation of an atomic mass is quite simple: it is a multiple of the mass of hydrogen.



Stanislao Cannizzaro (1826–1910) adopted the hydrogen atom as a standard of mass and set its atomic weight at 2. Others preferred a more massive standard in order to reduce experimental error.

A little while later, some technical difficulties were found. It turned out that H was not as reliable or simple a standard as was first thought. So for purely technical reasons, and to improve the accuracy of atomic weights, by 1850 it was decided that the oxygen atom, with mass 16.0 amu, would be the standard of reference. Then, 1 amu was 1/16 of the mass of a oxygen atom. Oxygen was chosen because it forms chemical compounds with many other elements, simplifying determination of their atomic weights. Sixteen was chosen because it was the lowest whole number that could be assigned to oxygen and still have an atomic weight for hydrogen that was not less than 1.

Only, this had a problem as well. The discovery of the neutron (in 1932 by James Chadwick) and of isotopes complicated the picture. In nature, pure oxygen is composed of a mixture of isotopes each with a different atomic mass. This was no problem for the chemists' calculations as long as the relative abundance of the isotopes in their reagents remained constant (a likely occurrence). Physicists, however, dealing with atoms and not molecules, required a unit that distinguished between isotopes. As early as 1927 physicists were using an atomic mass unit defined as equal to one sixteenth of the mass of the oxygen-16 atom (the isotope of oxygen containing 8 protons and 8 neutrons).

Thus the two amu scales were inconsistent: for chemists, 1 *u* was one-sixteenth of the average mass of the oxygen atoms in the chemist’s laboratory, and for physicists 1 *u* was one-sixteenth of the mass of a particular isotope of oxygen. In the years 1959-1961 the chemists and physicists reconciled this difference by agreeing to use the carbon-12 isotope as the standard, setting its atomic mass at exactly 12 *u*. The definition was ratified by the International Union of Pure and Applied Physics (IUPAP) in 1960 and the International Union of Pure and Applied Chemistry (IUPAC) in 1961, resolving finally the longstanding difference between chemists’ and physicists’ atomic mass scale.

Although not formally an SI unit, the atomic mass unit (formerly amu, now *u*) is accepted by the General Conference on Weights and Measures (CGPM) for use with SI.

The mass of the atomic mass unit is determined experimentally. According to the 1998 Committee on Data for Science and Technology (CODATA) recommendations, 1 *u* is $1.66053873 \times 10^{-27}$ kg, with a one-standard-deviation uncertainty of $\pm 0.000\,000\,13 \times 10^{-27}$ kg (relative uncertainty, 7.9×10^{-8}).

Avogadro's Number and the Mole

1. What number of iron atoms, each weighing 55.847 u, is necessary to get 55.847 g of Fe?

2. What quantity (in moles) of atoms of titanium are in 53.99 g of Ti? How many atoms is this?

3. (On-your-own problem) At \$450/oz (1 oz = 32 g), how much is 1.0 million atoms (1.0×10^6 atoms) of gold worth (in dollars)?

4. (Another take-home problem) Show that since 6.022×10^{23} atoms is 1 mol, that 6.022×10^{23} u is 1.00 g

5. How many atoms of iron are in 0.0255 mol of Fe? What mass, in g, is represented by 0.0255 mole of iron?

6. One molecule of CO_2 has a mass of 44.010 u. How many moles of CO_2 are in 15.01 g of the gas? How many molecules is this?

7. How many atoms of hydrogen are contained in 0.123 g of water?

8. (Another take-home...yes, another one) How many nitrogen atoms are in 1.50 g of the fertilizer ammonium phosphate, $(NH_4)_3PO_4$?

(The answer is 1.82×10^{22} atoms of N, I think)

Chapter 3

Chemical Compounds
and
Formula Stoichiometry

CHEMICAL COMPOUNDS

CHAPTER 3

INTRODUCTION

Chemical formulas are important because they provide us much information about the chemical compound. For example, the chemical formula can be used to identify the compound by name, calculate the molar mass and percentage composition of the compound, and mass and mole relationships of the various atoms in the compound. We learned in the laboratory how to name compounds and, in lecture, many features of the periodic table. Additionally, this chapter introduced the differences between ionic and molecular compounds.

GOALS

1. Many features of the periodic table have by now been introduced. These are important in the overall understanding of elements and molecules.
2. It is important to understand the difference between ionic and molecular compounds. As well, it is important to understand the differences between the different forms of that the "formula" can take (molecular, empirical, and structural).
3. Calculations involving the mole in regards to compounds is vital for future chapters. Percentage composition is important both experimentally (to determine the chemical formula) and theoretically.
4. The basics of nomenclature have been covered in lab. Oxidation states of simple monatomic ions is simply the ionic charge. More on oxidation states will be covered as we proceed through the book. For now, simply understanding **Section *3-5*** and **Example 3-7** will suffice.

DEFINITIONS

You should have a working knowledge of at least these terms and any others used in lecture.

Group
Family
Period
Long period
Short period
Metal
Nonmetal
Metalloid
Semiconductor
Semimetal
Element
Atom
Compound
Molecule
Anion
Cation
Binary compound
Ionic compound
Molecular compound
Molecular mass
Molar mass
Percentage composition
Law of Multiple Proportions
Empirical formula
Molecular formula
Structural formula
Oxidation state
Oxidation number
Salt
Acid
Hydrate

Formula Stoichiometry: Percentage Composition

1. What is the percent composition of CO_2?

2. What is the % composition of $(NH_4)_3PO_4$ (149.09 g/mol)?

   Just the %N calculation is shown here

3. By some means the percentage composition of a compound containing only Ti and O was determined to be 40.06% O and 59.94% Ti. What is the empirical formula of this compound?

   Start by assuming that there is 100g of Ti_xO_y. The mass doesn't matter; it's a convenience.

4. The percent composition of an ionic compound containing only Na, S, and O is 32.37% Na and 22.57% S. What is the empirical formula of the compound?

5. A compound containing only carbon, hydrogen, an oxygen was analyzed and found to be 40.00%C and 53.28%O. The molar mass of the compound was independently determined to be 180.2 g/mol. What is the molecular formula of the compound?

An Example of Percentage Composition

Hydrated sodium stannate has the percentage composition of 44.50% Sn, 17.24% Na, 35.99% O, and 2.27% H. Determine what the correct formula is for hydrated sodium stannate. The correct order of elements is

$$\mathrm{Na}_a\mathrm{Sn}_b\mathrm{O}_c \bullet d\ \mathrm{H_2O}$$

Experimental Determination of a Chemical Formula

To experimentally determine the formula of the compound formed from the reaction of magnesium and bromine, a student added 1.174 g of magnesium to a solution of hexane containing 1.142 g of bromine, Br_2. The bromine solution turned completely colorless but some magnesium metal remained. The magnesium metal was collected by filtration, washed, and dried. The mass of magnesium remaining was 1.001 g. From these data, calculate the experimentally determined formula for the magnesium-bromine compound.

Percentages in Calculations

1. When 0.100 g of magnesium reacts in air, 0.266 g of magnesium oxide is formed. What is the percentage by mass of the magnesium and oxygen in magnesium oxide?

2. The three naturally occurring isotopes of potassium are ${}^{39}_{19}\mathrm{K}$ (38.963707 u, 93.2581%), ${}^{40}_{19}\mathrm{K}$ (39.963999 u), and ${}^{41}_{19}\mathrm{K}$ (6.7302%) as measured in a mass spectrometer. What is the isotopic mass of ${}^{41}_{19}\mathrm{K}$?

3. Determine the mass percentage of water in $CuSO_4{\cdot}5\ H_2O$ (copper(II) sulfate pentahydrate).

4. Calcite, which gets its name from "*chalix*" the Greek word for lime, is one of the most common minerals on the face of the Earth, comprising about 4% by weight of the Earth's crust and is formed in many different geological environments. The formula for this ubiquitous and interesting mineral is $CaCO_3$. Calcium carbonate becomes marble from the heat and pressure of metamorphic events. What number of ${}^{13}_{6}C$ (1.108% abundance) atoms is present in 250.0 g of calcite?

5. Your instructor uses a “Triple-15” fertilizer. This designation represents a fertilizer that contains 15% each of nitrogen (N), P_2O_5, and K_2O. What mass of potassium is delivered to the soil when 15 lb of fertilizer are distributed?

6. Glucose contains only carbon, hydrogen, and oxygen. It has a percentage composition of 40.00% C and 6.71% H. The molecular mass as determined by mass spectrometric analysis is 180.2 u. What is the empirical and molecular formula for glucose?

Mass, Moles, and Percentages in Calculations

1. Ibuprofen, the pain-relief medication in, for example, Advil™ contains only carbon, hydrogen, and oxygen. During combustion analysis, 21.74 mg of pure ibuprofen produced 60.29 mg CO_2 and 17.09 mg H_2O. What is the empirical formula of ibuprofen?

2. Ibuprofen is an acidic compound containing a single carboxylic acid group (-COOH). Carboxylic acids react with bases in the same way that all acids react with bases. When 55.6 mg of ibuprofen is dissolved in water, 5.34 mL of 0.0505 M NaOH is required to completely neutralize the solution. What is the molar mass of the compound? Based upon the results of question 1, what is the molecular formula of ibuprofen?

Extra Stoichiometry Study Questions

1. What quantity, in moles, of water is formed by the reaction of 0.112 mol of oxygen gas with excess hydrogen gas?

2. In the same reaction as above, what quantity in moles of hydrogen gas are consumed? (Just for fun... how many molecules of hydrogen gas were consumed?)

3. If 1.000 mol of hydrogen gas was available, how much remains?

4. The reaction of silver nitrate with potassium chloride yields silver chloride as one of the products. What mass, in grams, of silver chloride will be produced by the reaction of 1.00 g of silver nitrate with 0.439 g of potassium chloride?

5. Assume that gasoline is entirely C_7H_8 (toluene). What mass of CO_2 will be produced by burning 1 tank (48.0 L) of gasoline in your car? The density of toluene is 0.866 g/mL. (FYI, gasoline is an enormously complex mixture of aliphatic and aromatic hydrocarbons with the toluene being a major component.)

Chapters 4 and 5

Chemical Reactions
and
Reaction Stoichiometry

CHEMICAL REACTIONS: REACTION STOICHIOMETRY

CHAPTER 4

INTRODUCTION

Chemical reactions are common experiences for us yet we rarely think about the process itself or that quantities of reactants and products are involved. This chapter starts by introducing the *chemical equation*, a shorthand way of describing a chemical reaction including the mass (or mole) balance. Nomenclature plays an important role in writing chemical equations so earlier studies of chemical formulas and nomenclature tie directly into this aspect of chemistry. After writing a correct and balanced chemical equation, it is then possible for the scientist (that would be you) to perform calculations involving masses or moles of reactants and products. Thus, it is possible to calculate the mass of the product expected from a given amount reactant and *vice versa*. Often one reactant is in limiting quantity so some amount of it remains after the reaction is complete. Through stoichiometric calculations, it is possible to calculate the amount of material remaining in a reaction involving limiting reagents. Finally, since many reactions are performed using solutions, the concept of concentration (often in molarity) is introduced and used.

GOALS

1. You must be able to complete and balance a variety of chemical reaction types. It is especially important that you can write and balance a reaction from the names of the reactants and products. Remember, a complete reaction also has the states of matter associated with each reactant and product.
2. Mass-to-mass (mole-to-mole) calculations are at the heart of the sciences which use chemistry as their foundation. It is essential that you can handle virtually any stoichiometric calculation. Remember, stoichiometry revolves around the mole ratio and the conversion of the moles of known substance to the moles of the substance desired.
3. Stoichiometric calculations are no different for problems involving one reactant in limiting quantity but it is important that you are able to recognize and do calculations involving limiting reagents.
4. Molarity as a concentration unit is important in performing stoichiometric calculations in situations such as simple solution chemistry and titration. More concentrated solutions may be diluted. You should know how to perform this operation and determine the new concentration of the solution after dilution.
5. Sometimes not all of the expected product in a reaction is realized experimentally. The percentage yield is a measure of this difference.
6. Nomenclature is becoming increasingly more important, especially in writing reactions.

DEFINITIONS

You should have a working knowledge of at least these terms and any others used in lecture.

Chemical equation
Chemical reaction
Chemical balance
Limiting reagent
Titration

Mole ratio
Concentration
Molarity
Molar concentration
Dilution

Solute
Solvent
Solution
Theoretical yield
Actual yield
Percentage yield

REACTIONS IN AQUEOUS SOLUTION

CHAPTER 5

INTRODUCTION

We have been introduced to a wide array of chemical reactions and reaction types. Additionally, we have seen that substances can fall into 3 different classes of electrolytes; weak, strong, and non-electrolytes. These concepts together and others as well allow us to predict reaction products and determine the reagents necessary to form a desired product. Of particular importance is writing and balancing oxidation-reduction reactions. The instructor recommends the balancing technique outlined in lecture over those of the authors of the textbook, but the textbook method works equally well. Finally, while titration was discussed in the prior chapter and laboratory, the formal introduction came in chapter 5.

GOALS

1. You should be able to write and balance chemical reactions of *all* reaction types presented in lecture and lab from formulas *and* names. Nomenclature is especially important in this chapter.
2. You should be able to identify the species by type (oxidized, oxidizing agent, reduced, etc.) in oxidation-reduction reactions.
3. The concept that compounds can be weak and strong electrolytes or non-electrolytes is critical in writing and predicting reactions.
4. Knowledge of the activity series and its use is important for specific reaction types - single-displacement.
5. Understanding titration and the stoichiometric calculations involved is very important.

DEFINITIONS

You should have a working knowledge of at least these terms and any others used in lecture.

Acid
Strong acid
Weak acid
Base
Strong base
Weak base
Salt
Electrolyte
Strong electrolyte
Weak electrolyte
Nonelectrolyte
Combination reaction
Decomposition reaction
Single-displacement
Double-displacement
Metathesis
Combustion
Neutralization
Oxidation-reduction reaction
Oxidation
Reduction
Oxidizing agent
Reducing agent
Whole molecular equation
Whole ionic equation
Net ionic equation
Precipitate
Titration
Titrant
Burette
Volumetric flask
Pipette
Indicator
Equivalence point
Endpoint
Standardization

Reaction Stoichiometry I
Mass-Mole Relationships and Limiting Reactant

1. According to the equation given, what mass of H_2 can be produced by the reaction of 10.0 g of Zn with the hydrochloric acid?

$$Zn_{(s)} + 2\ HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$$

2. According to the equation of the reaction of H_2 with O_2, what mass of water can be made from the mass of hydrogen from question 1?

$$2\ H_{2(g)} + O_{2(g)} \rightarrow 2\ H_2O_{(g)}$$

3. For the equation…

$$AgNO_{3(aq)} + BaCl_{2(aq)} \rightarrow AgCl_{(s)} + Ba(NO_3)_{2(aq)}$$

   a) Balance the equation

   b) How many moles of silver chloride can be produced starting with 2.0 moles of $AgNO_3$? …1.5 moles $BaCl_2$?

   c) How many mol of AgCl will be produced if 3.00 g of $BaCl_2$ react with a sufficient amount of $AgNO_3$?

d. What mass of $BaCl_2$ is necessary to form 5.15 g of AgCl?

4. What mass of water would be formed from the reaction of 5.0 g H_2, and 30.0 g O_2? How much of the excess reagent remains unreacted?

$$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g)$$

5. Calculate the mass of iron(III) chloride produced when 1.05 g iron(II) chloride reacts with 1.10 g of potassium permanganate ($KMnO_4$) according the to the equation...

$$5\ FeCl_2 + KMnO_4 + 8\ HCl \rightarrow 5\ FeCl_3 + MnCl_2 + KCl + 4\ H_2O$$

Reaction Stoichiometry II
Extra Stoichiometry Study Questions

---

1. a. What quantity, in moles, of water are formed by the reaction of 0.112 mol of oxygen gas with excess hydrogen gas?

   b. In the same reaction as above, what quantity in moles of hydrogen gas are consumed? (Just for fun... how many molecules of hydrogen gas were consumed?)

   c. If 1.000 mol of hydrogen gas was available, how much remains?

2. The reaction of silver nitrate with potassium chloride yields silver chloride as one of the products. What mass, in grams, of silver chloride will be produced by the reaction of 1.00 g of silver nitrate with 0.439 g of potassium chloride?

3. Assume that gasoline is entirely C_7H_8 (toluene). What mass of CO_2 will be produced by burning 1 tank (48.0 L) of gasoline in your car? The density of toluene is 0.866 g/mL. (FYI, gasoline is an enormously complex mixture of aliphatic and aromatic hydrocarbons with the toluene being a major component.)

Reaction Stoichiometry
Parallel Reactions

Petrucci & Harwood, Chapter 4

The manufacture of ethyl alcohol, C_2H_5OH, yields diethyl ether, $(C_2H_5)_2O$, as a by-product. The complete combustion of a 1.005-g sample of the product of this process yields 1.963 g CO_2. What must be the mass percents of C_2H_5OH and of $(C_2H_5)_2O$ in this sample?

Reaction Stoichiometry III
Yield, Concentration, and Solution Stoichiometry

1. If 5.05 g of $CuCl_2$ reacts with excess tin, what mass of copper is expected from the reaction?

$$CuCl_{2(aq)} + Sn \rightarrow SnCl_{2(aq)} + Cu_{(s)}$$

   What is the percentage yield if only 2.20 g of Cu is recovered? Where did the lost product go?

2. What is the molar concentration of a solution prepared by dissolving 5.84 g of NaCl in sufficient water to make 1.0 L of solution?

3. What mass of NaOH is necessary to prepare 500.0 mL of 0.250 M NaOH? (molar mass of NaOH = 40.00 g/mol)

4. What quantity, in moles, of HCl are present in 10.00 mL of 0.125 M HCl? What would the new HCl concentration be if 10.00 mL of water were added to the original 10.00 mL of 0.125 M HCl?

5. What is the concentration of the solution prepared by diluting 10.00 mL of 0.125 M HCl to 40.00 mL?

6. How might you prepare 500.0 mL of 0.10 M HC1 from concentrated HCl (12 M)?

7. What volume (in mL) of 0.15 M HCl will react with 2.50 g of Na_2CO_3 according to the equation

$$2\ HCl_{(aq)} + Na_2CO_{3(s)} \rightarrow 2\ NaCl_{(aq)} + CO_2(g) + H_2O_{(l)}$$

8. What is the concentration of a $KMnO_4$ solution if 15.16 mL of the $KMnO_4$ solution is required to titrate to equivalence point 25.00 mL of 0.08765 M $FeCl_2$?

$$5\,FeCl_2 + KMnO_4 + 8\,HCl \rightarrow 5\,FeCl_3 + MnCl_2 + KCl + 4\,H_2O$$

9. What is the percentage purity of a sample of impure oxalic acid ($H_2C_2O_4$, 90.04 g/mol) if 23.42 mL of the $KMnO_4$ from problem 8 was required to titrate 1.5960 g of the solid?

Concentration Calculation Involving Parts-per-Million

A solution used to chlorinate a home swimming pool contains 7% chlorine by mass. An ideal chlorine level for the pool is one part per million (1 ppm). If you assume densities of 1.10 g/mL for the chlorine solution and 1.00g/mL for the swimming pool water, what volume of the chlorine solution, in liters, is required to produce a chlorine level of 1 ppm in a 18,000-gallon swimming pool?

Types of Chemical Reactions

Combination

$2\ H_{2(g)} + O_{2(g)} \rightarrow 2\ H_2O_{(g)}$
$S_{(s)} + O_{2(g)} \rightarrow SO_{2(g)}$
$SO_{3(g)} + H_2O_{(l)} \rightarrow H_2SO_{4(aq)}$

A special case of combination – Acid anhydride-Base Neutralization

$SO_{3(g)} + NaOH_{(aq)} \rightarrow NaHSO_{4(aq)}$

Decomposition

$2\ H_2O_{(g)} \xrightarrow{\Delta} 2\ H_{2(g)} + O_{2(g)}$
$Cu(OH)_{2(s)} \rightarrow CuO_{(s)} + H_2O_{(l)}$

Metathesis (Double-Displacement)

$AgNO_{3(aq)} + HCl_{(aq)} \rightarrow AgCl_{(s)} + HNO_{3(aq)}$
$Pb(NO_3)_{2(aq)} + 2\ KI_{(aq)} \rightarrow PbI_{2(s)} + 2\ KNO_{3(aq)}$
$CaCO_{3(s)} + 2\ HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2CO_{3(aq)}$
but H_2CO_3 undergoes ***decomposition*** quickly: $H_2CO_{3(aq)} \rightarrow H_2O_{(l)} + CO_{2(g)}$ thus,
$CaCO_{3(s)} + 2\ HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$

A special case of metathesis - Acid-Base Neutralization

$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$
$2\ HNO_{3(aq)} + Ba(OH)_{2(aq)} \rightarrow BaCl_{2(aq)} + 2\ H_2O_{(l)}$

Single-Displacement

$Zn_{(s)} + 2\ HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$
$2\ AgNO_{3(aq)} + Cu_{(s)} \rightarrow 2\ Ag_{(s)} + Cu(NO_3)_{2(aq)}$

Oxidation-Reduction

$MnO_4^- + 5\ Fe^{2+} + 8\ H^+ \rightarrow Mn^{2+} + 5\ Fe^{3+} + 4\ H_2O$

Combustion

$2\ C_2H_{6(g)} + 7\ O_{2(g)} \rightarrow 4\ CO_{2(g)} + 6\ H_2O_{(g)}$
$C_2H_6O_{(g)} + 3\ O_{2(g)} \rightarrow 2\ CO_{2(g)} + 3\ H_2O_{(g)}$

SOLUBILITY RULES FOR IONIC COMPOUNDS

The following table outlines the general rules for predicting the solubility and electrolyte nature of ionic compounds in aqueous solution. These rules are also helpful in predicting the outcome of some reactions of ionic compounds in aqueous solution.

In general, a chemical reaction involving ionic substances will occur if a precipitate, gas, or weak electrolyte is formed as a product. The rules given are for aqueous solution and for ionic compounds only. The rules are not exhaustive and exceptions exist. The rules are not applicable for nonaqueous solutions or molecular compounds.

In general...

1. Group IA compounds including NH_4^+	All compounds are soluble and strong electrolytes.
2. Compounds of NO_3^-, $CH_3CO_2^-$, ClO_4^-	All compounds are soluble and strong electrolytes.
3. Compounds of Cl^-, Br^-, I^-, CN^-, SCN^-	All compounds are soluble and strong electrolytes except those of Ag^+, Hg_2^{2+}, and Pb^{2+}.
4. Compounds of SO_4^{2-}	All compounds are soluble and strong electrolytes except those of Pb^{2+}, Hg_2^{2+}, and Group IIA (Ca to Ra). $CaSO_4$ and Ag_2SO_4 are sparingly soluble.
5. Compounds of OH^- and S^{2-}	All compounds are insoluble and weak or nonelectrolytes except those of Group IA and Group IIA (Ca to Ra). Group IIA sulfides, Al_2S_3, and Cr_2S_3 all decompose in aqueous solution.
6. Compounds of CO_3^{2-}, PO_4^{3-}, AsO_4^{3-}, CrO_4^{2-}	All compounds are insoluble and weak or nonelectrolytes except those of Group IA and NH_4^+.

The formation of carbonic acid in a reaction usually results in the spontaneous decomposition:

$$H_2CO_3(aq) \rightarrow CO_2(g) + H_2O(l)$$

The formation of sulfurous acid (not sulfuric acid) also produces an analogous reaction:

$$H_2SO_3(aq) \rightarrow SO_2(g) + H_2O(l)$$

For this reason, CO_2 and SO_2 are two examples of acidic anhydrides.

Chapter 6

The Gaseous State

THE GASEOUS STATE

CHAPTER 6

INTRODUCTION

The behavior of gases is an aspect of chemistry that we give little thought to in our everyday life, but in reality is critical to our survival. Air is a mixture of mostly oxygen and nitrogen, the behaviors of each independently and in the mixture can be modeled by the ideal gas law. Helium balloons float while a balloon with argon will drop quite rapidly. These behaviors can also be modeled with the ideal gas law. Moreover, we can use physics and kinetic molecular theory to develop equations relating basic physical principles to observable phenomena.

GOALS

1. You should be able to relate P, V, T, and n.
2. It is important to be able to do problems involving changes in P, V, T, and n as well as calculate a single observable parameter knowing the others.
3. The ideal gas law can be used in its native form and modified in a variety of ways to solve fundamental problems involving gases such as the molar mass of the gas and the density of a gas.
4. Dalton's law of partial pressures deals with gases in a more experimental and practical way. You must be able to work with pure gases as well as mixtures of gases.
5. Kinetic molecular theory helps us to understand the relationship of microscopic properties and macroscopic properties.
6. You will not be held responsible for non-ideal behavior or Graham's law on the main portion of any exam. These are topics that might be covered for extra credit.

DEFINITIONS

You should have a working knowledge of at least these terms and any others used in lecture.

Pressure and units
Barometer
Barometric pressure
Manometer
Open-end manometer
Closed-end manometer
Boyle's law
Charles' law
Gay-Lussac's law
Ideal gas law
Universal gas constant
Dalton's law
Kelvin temperature
Mole fraction
Partial pressure
Standard temperature and pressure
Kinetic-molecular theory
Root-mean-square velocity
Kinetic energy
Diffusion
Effusion

Gas Laws I
The Ideal Gas Law

1. What quantity, in moles, of helium are in a 3.0 L Mickey Mouse balloon at Disneyland if the pressure in the balloon is 754 torr and the temperature is 24.2°C?

2. When measured at STP, what volume will 0.35 moles of oxygen gas occupy?

   When warmed to room temperature (25.0°C) and maintained at standard pressure, what will be the new volume?

3. Using the following data, determine the molar mass of the unknown compound.

   The mass of a 255.5 mL flask is 55.144 g (with the mass of the air subtracted out). After filling the flask with an unknown gas at the laboratory barometric pressure, the mass of the flask and gas was measured to be 55.363 g. The barometric pressure was determined to be 742.5 mm Hg using a mercury barometer. The laboratory temperature was 22.4°C.

   What is the molar mass of the gas?

4. A mountaineer blows up a balloon to 3.00 L at sea level where the pressure is 754 mm Hg. He flies to Tibet with his balloon and runs up to 20,000 feet on the way to the top of Mt. Everest, where the pressure is 371 mm Hg. What is the volume of the balloon? (Assume the temperature didn't change.)

5. Your ears are essentially a closed air-space inside your head. Mostly surrounded by bone, there is only one flexible wall enclosing this air-space – the ear drum. The volume inside the air space is not large at only about 1-2 mL. Assuming a volume of 1.0 mL and that you feel pain due to pressure on the ear drum when the volume of the inner ear is reduced by 0.05 mL, what pressure over atmospheric (1.0 atm) is necessary on the ear drum to cause pain? (Report the answer in atm, torr, psi, and Pa)

6. A mountaineer blows up a balloon to 3.00 L at sea level where the pressure is 754 mm Hg and the temperature is 22.0°C. He flies to Tibet with his balloon and runs up to 20,000 feet on the way to the top of Mt. Everest where the pressure is 371 mm Hg and the temperature is -15.5°C. What is the volume of the balloon?

Gas Laws I
The Ideal Gas Law
Additional Problems

1. In an experiment involving Boyle's law, a graph of which data produces a straight line; *P* vs. *V* or *P* vs. 1/V?

2. When a high-diver enters the water after jumping from a diving board, she may descend at much as 10 feet (or more) deep into the pool. At 10 feet, the total pressure on the diver will be 1.3 atm. Assuming that the diver took a deep breath (7.0 L of air in her lungs) at 1.0 atm just before entering the water, what will be the volume of air in her lungs when she gets to 10 feet of depth? Assume that the diver's air spaces (lungs) are completely flexible and she did not exhale while descending.

3. In an experiment involving Amonton's law, a graph of which data produces a straight line; *P* vs. *T* or *P* vs. 1/T?

4. Scuba tanks are given a "Maximum Working Pressure" (MWP) rating (which represents the highest pressure the tank can safely hold at a given temperature) and a “Hydrostatic Test Pressure” (HTP) rating (which represents the highest pressure the tank can hold before it may rupture, and is five-thirds the MWP). In the U.S., the MWP for a tank with an international yoke valve is 206.8 bar at 20.0°C. If the tank is filled to 200.0 bar at 20.0°C, what is the maximum pressure that the tank reaches if left in the sun on a boat, if the tank temperature increases to 39.0°C?

   Does the pressure exceed either (or both) the MWP or HTP? Will the scuba tank rupture?

5. A football is inflated to a pressure of 1.00×10^3 torr in a room at 25°C. If the game is played at 10°C, what will the pressure in the ball be, neglecting any volume change in the ball and assuming that it doesn't leak?

Gas Laws II
Gases in Reaction Stoichiometry

1. What volume of water vapor (measured at 755 mm Hg and 200.0°C) will be produced by the reaction of 32.0 g of oxygen gas with plenty of hydrogen gas?

2. What volume of oxygen gas, measured at 25.0°C and 1.00 atm would be formed by the complete decomposition of 2.00 g of $KClO_3$ (122.55 g/mol) at 250.0°C? (The products of the decomposition reaction are potassium chloride and oxygen gas.)

3. 3.0 L of nitrogen gas are reacted at 250.0°C and 1.5 atm with excess hydrogen gas to form ammonia. What will be the volume of ammonia present when collected and measured at STP assuming the reaction has a 100% yield?

4. Assume for the moment that when green plants inspire CO_2 gas during photosynthesis they produce only glucose for energy storage. Also assume that the photosynthetic reaction by the plant is

$$6\ CO_2(g) + 6\ H_2O(g) \rightarrow C_6H_{12}O_6(s) + 6\ O_2(g)$$

   If over a period of time the glucose content of a plant increases by 425 g, what volume of CO_2 was removed from the air? All measurements were made at 22.0°C and 751.5 torr.

Gas Law II
Gases in Reaction Stoichiometry
Additional Problems

1. The amount of NO_2 on a very smoggy day in Houston, TX was measured to be 0.78 ppm (by mass). The barometric pressure was 1011 mbar. Calculate the partial pressure of the NO_2.

   This problem is a little tougher since you have to calculate the molar mass of air (which also means you have to look up the composition of standard air.

2. To minimize the possibility of loud and possibly dangerous “backfire”, welders using large oxygen-acetylene cutting torches must be careful to prevent the ratio of acetylene (C_2H_2) to oxygen from becoming perfectly stoichiometric. What is the ratio of the pressures of acetylene and oxygen when they are stoichiometrically mixed?

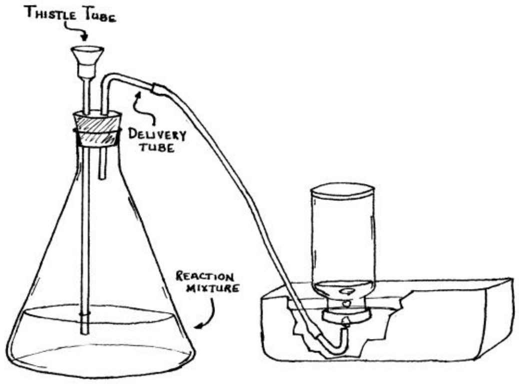
3. Sodium metal reacts with molecular chlorine gas to produce sodium chloride (and a great deal of energy). A closed glass container, fitted with a pressure gauge, has a volume of 3000 mL (±10 mL) and contains chlorine gas at 24.0°C and 1255 torr. On the bottom of the reaction vessel is a 6.90 g lump of sodium metal. The sodium is heated briefly to start the reaction, which continues to completion with no further intervention by the experimenter. The reaction vessel rises in temperature to 47.0°C. Predict the final pressure of the chlorine gas?

Gas Laws III
Dalton's Law of Partial Pressures in Applied Calculations

1. Air is about 78%$^{v}/_{v}$ N_2 and 21%$^{v}/_{v}$ O_2 (~1%$^{v}/_{v}$ Ar, CO_2, etc. total). Calculate the partial pressure (in mm Hg) of the oxygen gas and nitrogen gas when the total barometric pressure is 1003 mbar. (Recall that 1 atm = 1.01325 bar)

2. A scuba diver will often descend to 16 feet in the ocean where the pressure of the air being breathed is 1.5 atm. What is the partial pressure of the oxygen at this depth?

3. Refer to the gas collection device pictured. What is the partial pressure of the hydrogen gas in the gas collection bottle obtained when the water in the bottle is displaced by hydrogen from the reaction of zinc with hydrochloric acid? The volume of the bottle was measured to be 165 mL, the temperature of the equipment is assumed to be room temperature (23.2°C), and the barometric pressure is 751.9 mm Hg. It may be necessary to estimate or use an approximate value for the vapor pressure of water at this temperature.



4. What is the volume of the H_2 produced if the water vapor is removed?

5. Nitroglycerin explodes according to the equation

$$4\ C_3H_5N_3O_{9(l)} \rightarrow 12\ CO_{2(g)} + 10\ H_2O_{(g)} + 6\ N_{2(g)} + O_{2(g)}$$

What is the total pressure in a 1.0 L closed rigid container (perhaps a hole in the rock in a mine) when 200.0 g of nitroglycerine explodes. Assume for the problem that the temperature of the produced gases are 850°C.

Vapor Pressure of Water
at Several Temperatures

Temperature (°C)	**Vapor Pressure (mm Hg)**
15.0	12.79
16.0	13.63
17.0	14.53
18.0	15.48
19.0	16.48
20.0	17.54
21.0	18.65
22.0	19.83
23.0	21.07
24.0	22.39
25.0	23.76
26.0	25.21
27.0	26.74
28.0	28.35
29.0	30.04
30.0	31.82

Gas Laws III
Dalton's Law of Partial Pressures
Additional Problems

1. The amount of NO_2 on a very smoggy day in Houston, TX was measured to be 0.78 ppmv (parts-per-million by volume). The barometric pressure was 1011 mbar. Calculate the partial pressure of the NO_2.

2. A mixture of cyclopropane gas (C_3H_6) and oxygen gas in a 1.00:4.00 mol ratio is uncommonly used as an anesthetic gas. What mass of each gas is present in a 2.00 L steel container pressurized to 150.0 bar at 25.0°C?

3. “Mixed-air” divers often use standard air (78%N_2, 21%O_2, 1%Ar) which has been enriched to 32%O_2. As the dive tender aboard a marine science research vessel, it is your responsibility to fill scuba tanks with the proper air mix. A 12.5 L (internal volume) scuba tank is pressurized to 2550 psi with standard air. You add pure oxygen to the tank. What must the final pressure be so that the air has a composition of 32%O_2? All measurements are made at 25.0°C.

Mean Free Path in the Kinetic Molecular Theory of Gases

In the kinetic molecular theory of gases the mean free path of particles is the average distance the particle travels between collisions with other moving particles. If the velocities of the particles behave according to the Maxwell distribution, the following relationship applies:

$$l = \frac{kT}{\sqrt{2}\pi d^2 P}$$

where k is the Boltzmann constant (1.381×10^{-23} J/K), T is Kelvin temperature, P is pressure (in Pa), and d is the diameter (in m) of the gas particles.

The table lists some typical mean free path values for dry air (average molar mass 28.97 g/mol, average $d = 3.67 \times 10^{-10}$ m) at different pressures.

Vacuum range	**Pressure in kPa**	**Pressure (torr)**	**Molecules / cm^3**	**Mean free path**
Ambient pressure	101.3	760	2.7×10^{19}	68 nm
Low vacuum	30 – 0.1	225 – 0.75	$10^{19} – 10^{16}$	0.1 – 100 μm
Medium vacuum	0.1 – 0.01	0.75 – 0.075	$10^{16} – 10^{13}$	0.1 – 100 mm
High vacuum	$10^{-2} – 10^{-6}$	$0.075 – 7.5 \times 10^{-6}$	$10^{13} – 10^{9}$	10 cm – 1 km
Ultra high vacuum	$10^{-6} – 10^{-11}$	$7.5 \times 10^{-6} – 7.5 \times 10^{-11}$	$10^{9} – 10^{4}$	1 km – 10^5 km
Extremely high vacuum	$<10^{-11}$	$<7.5 \times 10^{-11}$	$<10^4$	$>10^5$ km

Gas Laws IV
Kinetic-Molecular Theory

1. What is the root-mean-squared velocity of nitrogen gas at 25.0°C?

2. What is average kinetic energy of the nitrogen molecules at 25.0°C?

3. What is the average kinetic energy (in J/mol) of SF_6 at 25.0°C? What is the average velocity of SF_6 at this temperature?

4. In a 1-meter-long time-of-flight (TOF) mass spectrometer, a helium ion traversed from the ion source to the ion detector in 12.25 ms. A gas of unknown identity traversed the TOF tube in 188 ms. What is the molar mass of the unknown gas?

5. The escape velocity of a object from Earth's gravitational field is about 25,000 mi/h. In units of miles-per-hour, what is the average velocity of helium at 0°C?

6. At what temperature does the velocity of a helium atom exceed the escape velocity of the Earth?

Chapter 7

Thermochemistry

THERMOCHEMISTRY

CHAPTER 7

INTRODUCTION

Thermochemistry is a facet of chemistry which combines reaction writing, completion, and balancing with the heat or energy absorbed or released in the chemical reaction. In essence, you might say that the energy in a reaction is a product or reactant normally hidden from the person writing the reaction. This chapter introduces a way to determine both theoretically (Hess's Law) and experimentally (calorimetry) the energy change (enthalpy) in a chemical reaction.

GOALS

1. You must have a working knowledge of all of the terms involved in thermochemistry.
2. You should be able to calculate the amount of heat transferred in a chemical reaction when experimentally measured in a calorimeter.
3. Hess's Law allows for the calculation of ΔH of a reaction without performing an experiment. You should be able to do Hess's law calculations.

DEFINITIONS

You should have a working knowledge of at least these terms and any others used in lecture.

Energy
Heat
Enthalpy
Endothermic
Exothermic
Calorimetry
Calorimeter
Specific heat
Heat capacity
Open system
Closed system
Isolated system
Heat of reaction
Enthalpy of reaction
System
Surroundings
Standard state
Work
Joule
First law of thermodynamics

Thermochemistry I: Energy Transfer and Calorimetry

1. What amount of work (in J) is performed on the surroundings when a 1.0 L balloon at 745 mm Hg at 25°C is heated to 45°C? (1 L·atm = 101.325 J)

2. What quantity of heat (in J) is necessary to raise 3.00 L of water (*d*=1.00 g/mL) from 22.0°C to 63.0°C?

3. A 200.0 mL quantity of 0.40 M HCl was added to 200.0 mL of 0.40 M NaOH in a solution (constant pressure) calorimeter. The temperature of each solution was 25.10°C before mixing. After mixing the solution rose to a temperature of 26.60°C before beginning to cool. The heat capacity of the calorimeter was determined by separate experiment to be 55 J/°C. What is ΔH_{rxn} per mol of H_2O formed? Assume the solutions have a density of 1.00 g/mL and their specific heats are similar to water; c = 4.18 J/g·°C.

4. A 1.00 g sample of table sugar (sucrose, $C_{12}H_{22}O_{11}$) was burned in a bomb calorimeter (constant volume calorimeter) containing 1.50 kg of water. The temperature of the water in the calorimeter rose from 25.00°C to 27.32°C. What is the $\Delta H_{\text{combustion}}$ of sucrose in kJ/g and kJ/mol? The heat capacity of the calorimeter was determined by separate experiment to be 837 J/°C.

5. Camphor ($C_{10}H_{16}O$) has a $\Delta H_{\text{combustion}}$ of-5903.6 kJ/mol. A 0.7610 g sample of camphor was burned in a bomb calorimeter containing 2.00×10^3 g of water. The temperature of the water increased from 22.78°C to 25.06°C. What is the heat capacity of the calorimeter?

Determination of the Specific Heat of Copper Metal

Data

Copper

M_{Cu} = ____________

T_i = ____________

T_f = ____________

ΔT_{Cu} = ____________

Water in Calorimeter

V_{H_2O} = ____________

m_{H_2O} = ____________

c_{H_2O} = ____________

T_i = ____________

T_f = ____________

ΔT_{H_2O} = ____________

Calorimeter

C_{cal} = ____________

T_i = ____________

T_f = ____________

ΔT_{cal} = ____________

$$q_{Cu} + q_{H_2O} + q_{cal} = 0$$

Thermochemistry II: Calorimetry, Enthalpy, and Hess' Law

1. When 100.0 mL of 1.00 M HCl is mixed with 100.0 mL of 1.00 M NaOH, both initially at 21.1°C, are mixed in a two-cup calorimeter the temperature of the mixture rises to 27.9°C. Determine the ΔH of neutralization for the reaction

   $$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

   By a prior experiment, the heat capacity of the calorimeter was determined to be 125 J/°C. Assume the density of the final solution is 1.0 g/mL and the specific heat of the mixture is 4.18 J/g°C.

2. Consider the reaction

   $$C_{12}H_{22}O_{11(s)} + 12\ O_{2(g)} \rightarrow 12\ CO_{2(g)} + 11\ H_2O_{(g)}$$

   which has a ΔH of -5.65×10^3 kJ/mol ($C_{12}H_{22}O_{11}$). How much heat (energy) can be produced during the complete combustion of 100.0 g of sucrose?

3. If all of the energy in question 2 were used to heat 1.0 L of water at 22.0°C, what would the final temperature of the water be? (Assume 100% energy transfer to the water.)

4. Using standard enthalpies of reaction, calculate the $\Delta H°$ for the following reactions:

$$\mathbf{C_2H_{2(g)} + 2\,H_{2(g)} \rightarrow C_2H_{6(g)}}$$

$$\mathbf{2\,CH_{4(g)} + \tfrac{1}{2}\,O_{2(g)} \rightarrow C_2H_{6(g)} + H_2O_{(l)}}$$

TABLE OF SOME STANDARD ENTHALPIES OF FORMATION AND REACTION

$C_{(s)} + 2\ H_{2(g)} \rightarrow CH_{4(g)}$	$\Delta H^{\circ}_{f} = -74.9$ kJ/mol
$2\ C_{(s)} + 3\ H_{2(g)} \rightarrow C_2H_{6(g)}$	$\Delta H^{\circ}_{f} = -84.7$ kJ/mol
$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$	$\Delta H^{\circ}_{f} = -393.5$ kJ/mol
$CH_{4(g)} + 2\ O_2 \rightarrow CO_{2(g)} + 2\ H_2O_{(l)}$	$\Delta H^{\circ}_{combustion} = -890.4$ kJ/mol(CH_4)
$C_2H_{2(g)} + \frac{5}{2}\ O_{2(g)} \rightarrow 2\ CO_{2(g)} + H_2O_{(l)}$	$\Delta H^{\circ}_{combustion} = -1299.4$ kJ/mol(C_2H_2)
$H_{2(g)} + \frac{1}{2}\ O_{2(g)} \rightarrow H_2O_{(g)}$	$\Delta H^{\circ}_{f} = -241.8$ kJ/mol
$H_{2(g)} + \frac{1}{2}\ O_{2(g)} \rightarrow H_2O_{(l)}$	$\Delta H^{\circ}_{f} = -285.8$ kJ/mol
$Na_{(s)} \rightarrow Na_{(g)}$	$\Delta H^{\circ}_{sublimation} = 108$ kJ/mol
$Na_{(g)} \rightarrow Na^{+}_{(g)} + e^{-}$	$\Delta H^{\circ}_{ionization} = 459.9$ kJ/mol
$2\ Na_{(s)} + \frac{1}{2}\ O_{2(g)} \rightarrow Na_2O_{(s)}$	$\Delta H^{\circ}_{f} = -99.8$ kJ/mol
$Na_{(s)} + \frac{1}{2}\ O_{2(g)} + \frac{1}{2}\ H_{2(g)} \rightarrow NaOH_{(s)}$	$\Delta H^{\circ}_{f} = -98.9$ kJ/mol

Chapter 8

Quantum Theory

ATOMIC STRUCTURE

CHAPTER 8

INTRODUCTION

Knowledge of the electronic structure of the atom explains many observed phenomena, ranging from the emission of light from a glowing gas to why metals form cations and nonmetals form anions. The development of quantum mechanics ushered in a new branch of physics and chemistry and the quantum mechanical model of the atom remains the model in use today. The current model has its roots in spectroscopy; the study of electromagnetic radiation and its interaction with matter. Using electromagnetic radiation, it is possible to probe the microscopic structure of atoms and molecules. Possibly the most unique property of electromagnetic radiation, however, is its property of behaving like a wave and a particle simultaneously - known as wave-particle duality. With this seemingly paradoxical observation of wave-particle duality, it is possible to explain both the line spectrum observed for excited gas-phase atoms and the photoelectric effect, as well as many other observed phenomena. An essential part of understanding many of the subtle aspects of chemistry require a basic understanding of light and quantum mechanics.

GOALS

1. It is important to be able to relate wavelength, frequency, velocity, and energy of light. Additionally, knowledge of the region from the complete electromagnetic radiation spectrum a particular wavelength is in is very important.
2. You must be able to describe the difference between a line spectrum and a continuous spectrum and be able to explain the source of the line spectrum of a gas-phase atom.
3. You should be able to trace the history of the model of the atom up to the quantum mechanical model in use today. You should be also be able to discuss the strengths and weaknesses of each model.
4. Evidence exists that particles, such as electrons, exhibit wave-like characteristics, a hypothesis proposed by de Broglie. You should be able to calculate the apparent wavelength of a moving particle using de Broglie relationship.
5 Bohr's model of the atom placed the electrons in planetary orbits around the nucleus. Bohr's model is very compelling since it is quite intuitive and calculations with the Rydberg equation yield nearly perfect correlation between experiment and theory for the hydrogen atom line spectrum. It is important to know when the model breaks down (any atom except hydrogen) and what this model led to.
6. The quantum mechanical model of the atom describes the electronic structure of the atom in terms of probability regions around the nucleus *(orbitals).* These probability regions are described by wave-functions which have quantized solutions called the quantum numbers (n, l, m_1). It is essential that you be able to assign the quantum numbers to each electron in an atom.
7 Just as in Bohr's model, the quantum mechanical model places electrons in progressively higher energy levels defined by the orbital's quantum numbers. You must be able to draw the atomic orbital energy level diagram through at least the 5*p* subshells. You must also be able to properly place the electrons in the orbital energy diagram (using Aufbau, Pauli exclusion, and Hund's rule) and write the correct electron configuration.

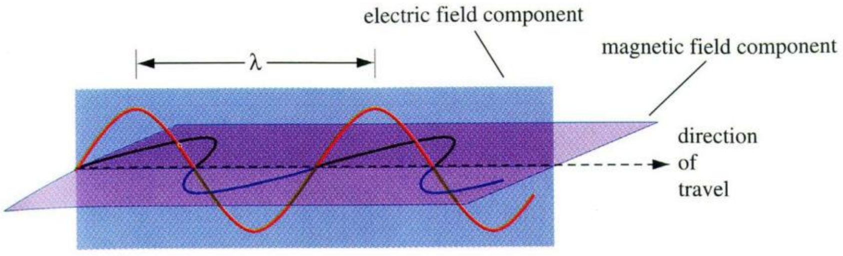
8. You should be able to correctly identify whether an atom is in an excited state based on the electron configuration or the atomic energy level diagram.

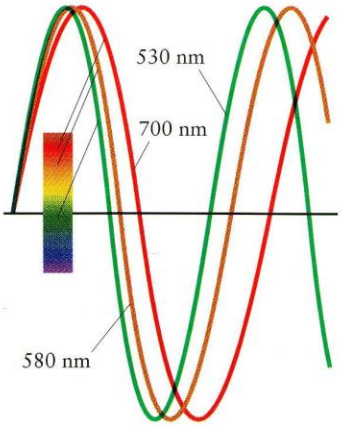
DEFINITIONS

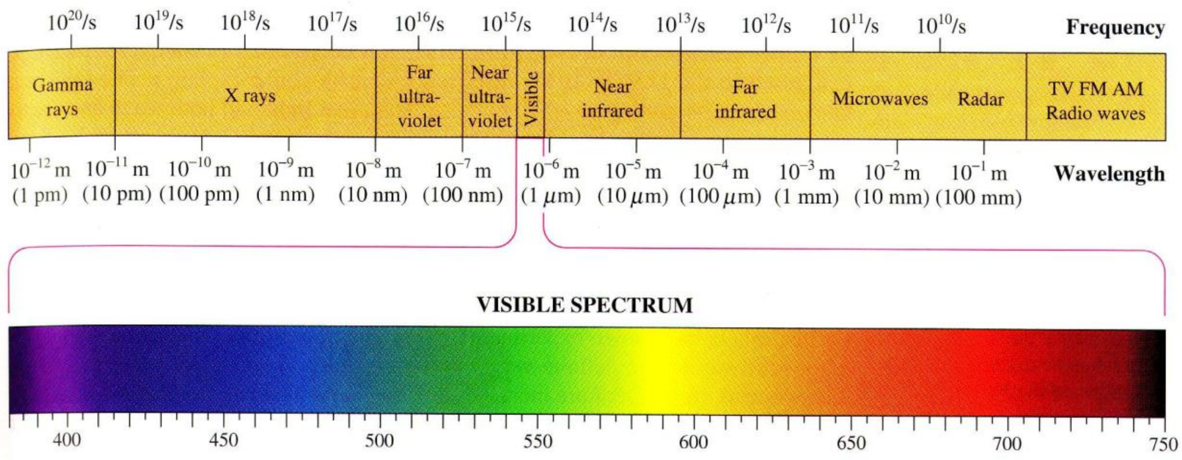
You should have a working knowledge of at least these terms and any others used in lecture

Light
Electromagnetic radiation
Angular-momentum quantum number
Spectrum
Line spectrum
Continuum
Continuous spectrum
Wavelength
Frequency
Speed of light
Hertz
Energy
Planck's constant
Photoelectric effect
Quantum
Photon
Wave
Rydberg's equation
Wave-particle duality
Ground state
Excited state
Orbit
Orbital
Principal quantum number
de Broglie relationship
Heisenberg uncertainty
quantum number
Magnetic quantum number
Spin quantum number *n*, *1*, m_l, m_s
s, *p*, *d*, *f*
Subshell
Aufbau
Pauli exclusion
Hund's rule
Degenerate energy
Atomic orbital energy-level diagram
Electron configuration

ELECTROMAGNETIC RADIATION







MEASURED QUANTIZED EVENT IN EACH REGION

γ-ray

x-ray

UV

Visible

IR

Microwave

Radiowave

Some Important Equations and Constants

Constants

$c = 3.00 \text{ x } 10^{8}\ {}^{\text{m}}/_{\text{s}}$

$h = 6.63 \text{ x } 10^{-34}\text{J}\cdot\text{s}$

$\sigma = 5.67 \text{ x } 10^{-8}\ {}^{\text{W}}\!/_{\text{m}^2\cdot\text{K}^4}$

Equations

$c = \lambda\nu$

Planck's Law: $E = h\nu = \dfrac{hc}{\lambda}$

de Broglie's Law: $\lambda = \dfrac{h}{m\text{v}}$ [v (velocity) not ν (frequency)]

Wien's Law: $\lambda_{\text{max}} = \dfrac{0.0029}{T}$

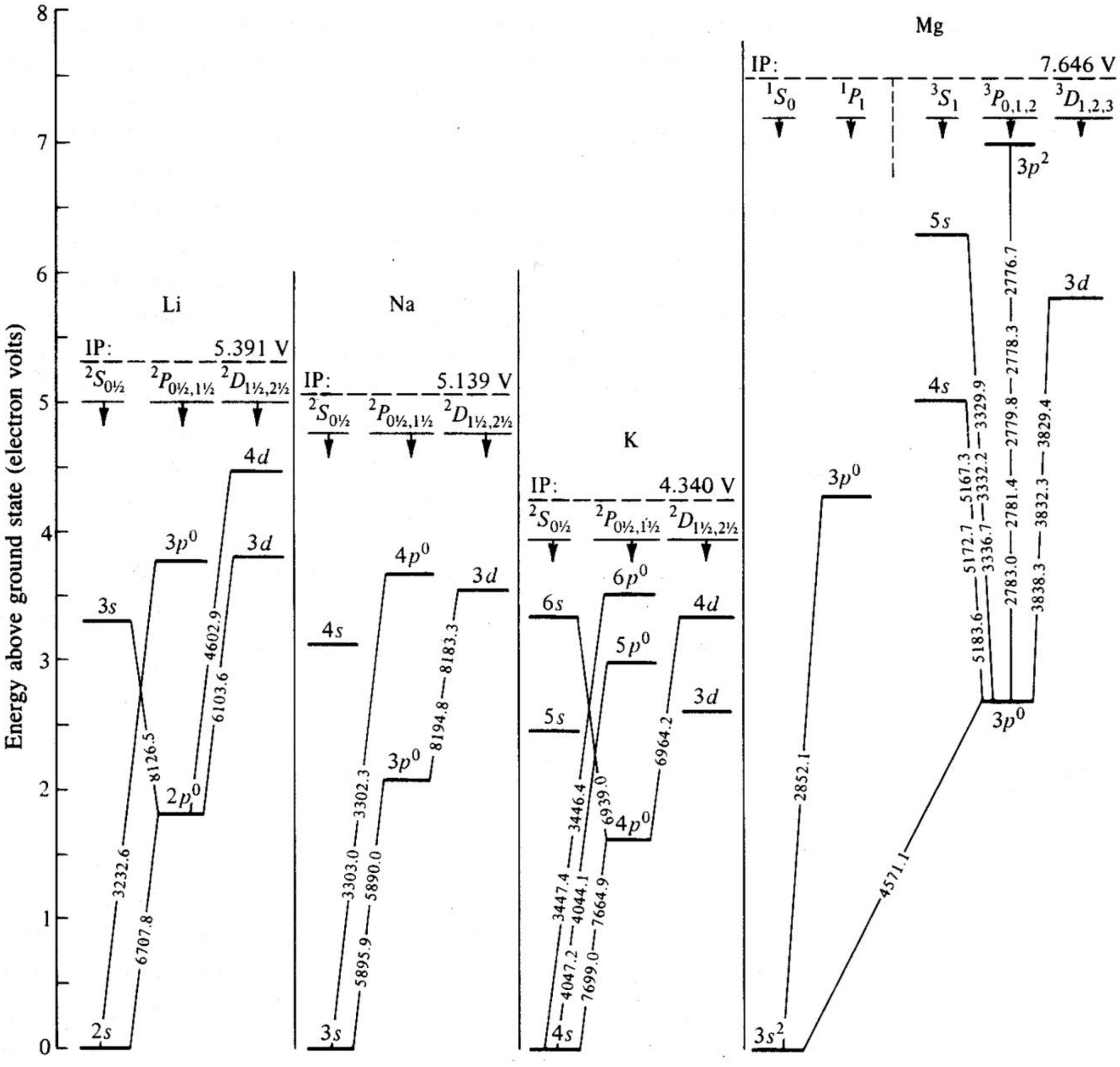
Stephan-Boltzmann Law: $F = \sigma T^4$

Inverse-Square Law of Light: $F = \dfrac{L}{4\pi d^2}$

Einstein's Energy/Mass Relationship: $E = mc^2$

Schwarzschild Radius: $R = \dfrac{2GM}{c^2}$

Atomic term diagram for Li, Na, K, and Mg. Wavelengths are given in angstroms. (IP = ionization potential in eV.)



Atomic Structure I: Electromagnetic Radiation

1. What are the colors of the visible light spectrum in order from short to long wavelength? What is the approximate wavelength of each color?

2. If the thermonuclear reactor in the center of our sun emits an x-ray photon, how fast will that photon be travelling?

3. What is the wavelength of a 650 nm photon in meters?

4. The “oldies” FM radio station, KRTH, transmits on 101.1 MHz. What is the wavelength of the radiation (in meters)?

5. The Sun emits the highest intensity of light at about 520 nm. What is the color and the energy (in J) of this radiation?

6. A particular star emits light with a λ_{max} of 485 nm. What is the surface temperature of the star? (Stellar emission can be considered to be blackbody.)

7. What is the kinetic energy of the electrons ejected from silver metal when it is struck by light with a wavelength of 200 nm?

8. The technique of photoelectron spectroscopy (an instrumental method that measures the kinetic energy of electrons produced photoelectrically from metallic targets) was used to identify the metals in an alloy. The electromagnetic radiation source was the Mg K_α x-ray emission at 0.9898 nm. The spectrum showed two photoelectron peaks; the first of which was at 2.0008×10^{-16} J and the other at 2.0014×10^{-16} J. The relative ratios of the two photoelectron emissions was 14:8. What is the composition of the alloy?

Work Functions for Photoelectric Effect

Element	**Work Function (eV)**	**Work Function (J)**
Aluminum	4.08	6.54×10^{-19}
Beryllium	5.0	8.01×10^{-19}
Cadmium	4.07	6.52×10^{-19}
Calcium	2.9	4.65×10^{-19}
Carbon	4.81	7.71×10^{-19}
Cesium	2.1	3.36×10^{-19}
Cobalt	5.0	8.01×10^{-19}
Copper	4.7	7.53×10^{-19}
Gold	5.1	8.17×10^{-19}
Iron	4.5	7.21×10^{-19}
Lead	4.14	6.63×10^{-19}
Magnesium	3.68	5.90×10^{-19}
Mercury	4.5	7.21×10^{-19}
Nickel	5.01	8.03×10^{-19}
Niobium	4.3	6.89×10^{-19}
Potassium	2.3	3.68×10^{-19}
Platinum	6.35	1.02×10^{-18}
Selenium	5.11	8.19×10^{-19}
Silver	4.73	7.58×10^{-19}
Sodium	2.28	3.65×10^{-19}
Uranium	3.6	5.77×10^{-19}
Zinc	4.3	6.89×10^{-19}

Adapted from *CRC Handbook of Chemistry and Physics*.

Atomic Structure II: Wave-Particle Duality

1. Using de Broglie's equation, calculate the apparent wavelength of an electron with a velocity of 40.2 m/s. ($m_{electron}$ = 9.11 x 10^{-28} g)

2. What is the "wavelength" of a baseball (0.181 kg) pitched at this speed? (FYI, 40.2 m/s = ~90 mi/hr)

3. What prevents us from actually measuring the wavelength calculated in problem 2?

4. Just for fun… What is the velocity of a rubidium atom when its de Broglie wavelength is equal to its atomic radius? The atomic radius of Rb is 248 pm. Approximately what temperature is this? Incidentally, this is called the Bose-Einstein condensate.

Atomic Structure III: Quantum Mechanics

1. Consider a hydrogen atom which has been electrically excited into emission. What is the energy of the light emitted when the electron falls from the $n = 4$ state to the $n = 2$ state?

2. What is the wavelength of the light produced in question 1? What is the color of the photon?

3. What region of the electromagnetic radiation spectrum is a photon produced by hydrogen atom in which the electron falls from $n = 7$ to $n = 6$?

4. The ionization energy of hydrogen atom is 2.179×10^{-18} J. When an electron recombines with an ionized hydrogen, it emits a continuum of radiation rather than a line spectrum. Explain this observation.

5. What is the wavelength of light emitted by a hydrogen atom when an electron at the ionization potential (*i.e.*, the electron is just beyond the $n = 7$ energy level) recombines with the hydrogen $n = 7$ energy level?

6. Explain the observation that, in the solar blackbody spectrum, dark lines are seen at exactly the same wavelengths as the emission lines for hydrogen.

7. What are the allowed quantum numbers for a ground-state electron in the outer-most orbital of magnesium metal?

9. Write the full electron configuration for ground-state

   nitrogen atom

   vanadium atom

   chromium atom

10. Write the abbreviated electron configuration for ground-state

    oxygen atom

    zirconium atom

    copper atom

11. What are one possible set of quantum numbers for an excited-state electron of magnesium metal in the which the electron has been promoted to a *d*-orbital in the fifth principle quantum shell?

12. When gaseous lithium metal is excited to emission in a flame, it produces a striking and beautiful red emission. One of the emission lines of lithium is 610.4 nm and is due to the excited state electron falling from a 3*d* orbital to a 2*p* orbital.

    What is the energy difference between the lithium 3*d* and 2*p* orbital?

    After the transition, is lithium at ground state?

Using the de Broglie Relationship

Using de Broglie's equation, calculate the apparent wavelength of an electron with a velocity of 40.2 m/s. ($m_{electron}$ = 9.11 x 10^{-28} g)

What is the "wavelength" of a baseball (0.181 kg) pitched at this speed? (FYI, 40.2 m/s = ~90 mi/hr)

For fun...
What is the velocity of a rubidium atom when its de Broglie wavelength is equal to its atomic radius? The atomic radius of Rb is 248 pm. Approximately what temperature is this? Incidentally, this is called the Bose-Einstein condensate

Chapter 9

The Periodic Table
and the
Periodic Law

PERIODIC PROPERTIES OF THE ELEMENTS

CHAPTER 9

INTRODUCTION

Quantum mechanics not only helps to describe and explain the microscopic structure of the atom, it also provides a theoretical basis to describe and explain macroscopic experimentally observed properties of the elements. Just a few of the periodic properties observed are atomic and ionic size, ionization energy, and electronegativity. Many observed chemical reactions can be predicted and explained simply by observing where the elements in the reaction are situated on the periodic table. The periodic table was discussed in some detail earlier in the semester; this chapter is thus mostly review and support for earlier discussion. From this chapter, it is clear that the table devised by Mendeleev without knowledge of quantum mechanics has much deeper theoretical roots than Mendeleev imagined.

GOALS

1. You should be comfortable with all terms, definitions, and historic figures discussed in the lecture and text. This is a very *descriptive* chapter - very few if any calculations are performed.
2. You must understand the periodic properties discussed in lecture and the text and to make predictions based on the properties.
3. The periodic table is divided into regions, blocks, families, periods, and element type. It is important to be able to distinguish each of these features. Study your notes from the earlier chapters.
4. The electronic structure of the atom determines its magnetic properties. You should be able to determine the electronic structure of the atom (chapter 9) and determine its magnetic properties.
5. It is important to know the exceptions to the orbital filling rules and have a basic understanding of why these exceptions exist.
6. Many reactions illustrating the periodic properties have been introduced that you should feel responsible for.

DEFINITIONS

You should have a working knowledge of at least these terms and any others used in lecture.

Refer to the review sheet for **Chapter 3** in addition to this study sheet.

Block
Periodic law
Law of Triads
Law of Octaves
Isoelectronic
Special stability
Actinides
Lanthanides
Paramagnetic
Diamagnetic
Atomic radius
Ionic radius
Ionization energy
Electron affinity
Metal oxide
Nonmetal oxide
Effective nuclear charge
Z_{eff}
Valence
Kernel
Special stability
Shielding
Electron configuration
Noble gas configuration