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 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 6.022×10^{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 $e/m = -1.76 \times 10^{8}\ ^{C}/_{g}$
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.

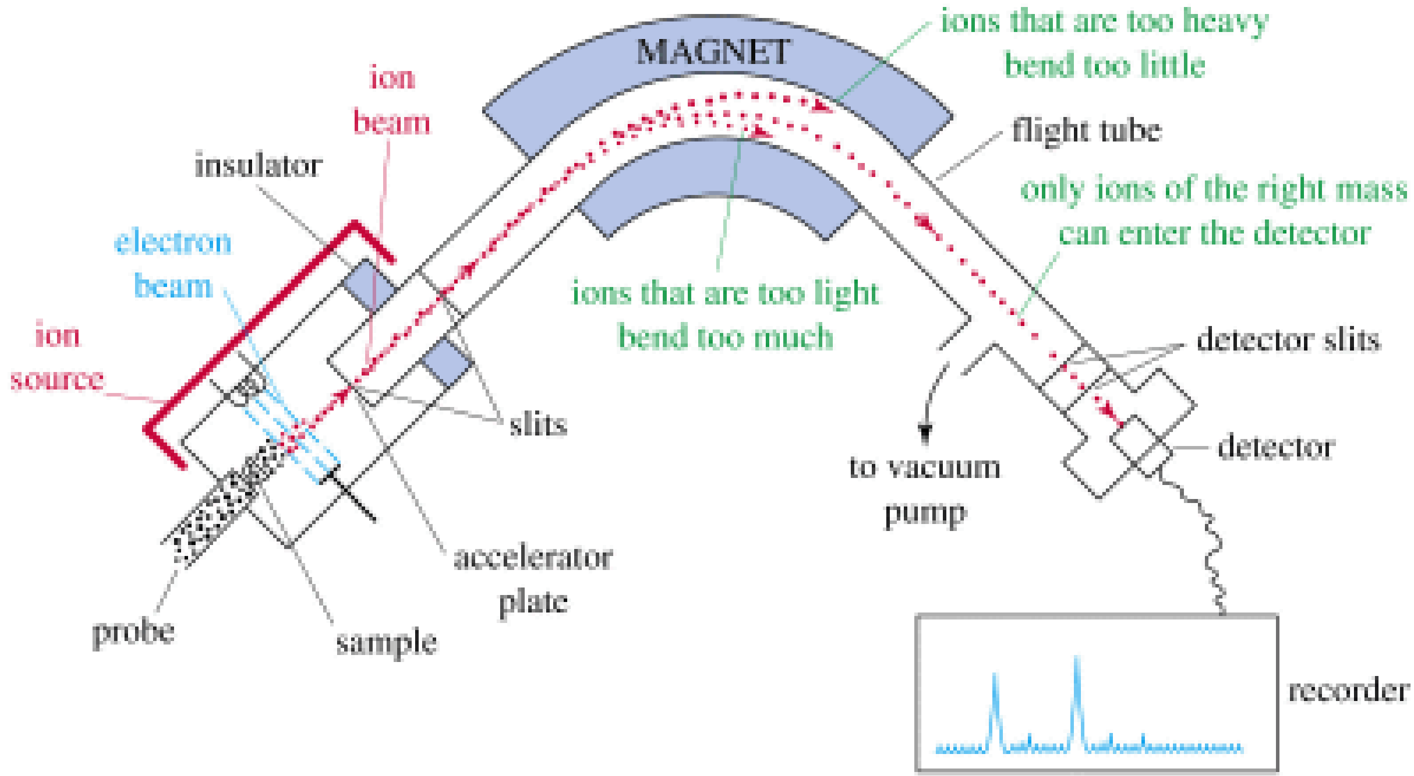
2. **All atoms of the same element are identical**
   Again, a good postulate for the time. It wasn't until 1932 that neutrons were discovered verifying the existence of nuclear isotopes. All atoms of the same element are the same or differ only by their number of neutrons.

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.

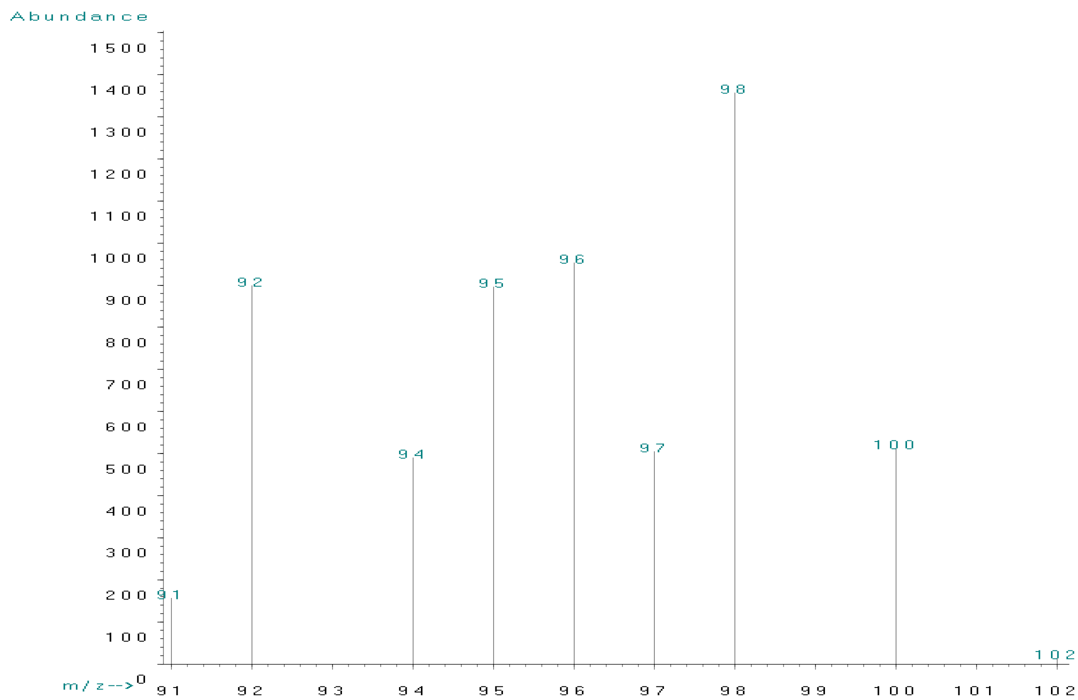
4. **The numbers and kinds of elements in a given compound are constant.**

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).

A Magnetic Sector Mass Spectrometer



ISOTOPIC ANALYSIS BY MASS SPECTROMETRY



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	

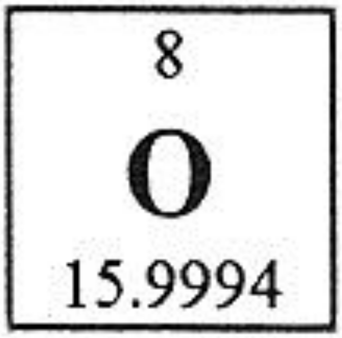
Average Mass of an Element

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

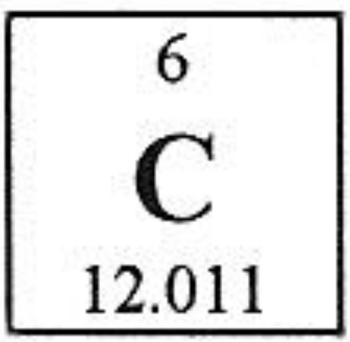
Brief and Mostly Accurate History of the AMU Scale

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 x 10^6 atoms) of gold worth (in dollars)?

4. (Another take-home problem) Show that since 6.022 x 10^{23} atoms is 1 mol, that 6.022 x 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 44.01 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)