**TLDR (Too Long, Didn’t Read) for Chapter 3- Atoms: The Building Blocks of Matter**

**TLDR 3.1 The Atom: From Philosophical Idea to Scientific Theory (pp. 68-73)**

Humans have been thinking about the nature of matter for a long time now. Democritus of Abdera (400 B.C.) was the first person (that we know of) whose musings on the properties of the atom were; 1) recorded and 2) somewhat on point. It only took another 2000 years or so before the next great thinker came along, and he was the first to really apply the scientific method to his claims. This man was John Dalton (1766-1844), an English schoolteacher who used experimental results to refine and develop his model of the atom.

1. List and then compare the main ideas about the atom proposed by Democritus and John Dalton.

2. Of Dalton’s atomic theory, which of the main points do we still consider valid? Which of his main ideas have had to change in light of more recent discoveries?

**TLDR 3.2 The Structure of the Atom (pp. 74-78)**

By the turn of the 20th century, a flood of new research was reported that revealed the fact that the atom was no longer “indivisible”, but was in fact, made up of one or more subatomic particles. The first of these to be “discovered” was the electron (why was the electron the first?), and in fact, electricity was being used to great effect for decades before its nature was determined. The first 30 years of the 20th century saw a monumental increase in our understanding of the structure and behavior of the atom, as well as the laws governing the behavior inside the nucleus, ushering in the “atomic age” and leading to the advent of nuclear power and the growing threat of nuclear holocaust. This section of your text reads like a history textbook leading you through some of the major discoveries and discoverers of this time. Do not bother memorizing particular dates, I am more concerned with you understanding the general direction of the advances in this field and the obvious pattern of discovery of each of the main subatomic particles based on their physical properties. You should be able to name the main discoverers and provide a cartoon depiction of the experiments that led to each discovery.

1. Draw a cartoon depiction of a cathode ray tube (also known at the time as a “Crookes Tube”). What experiment led scientists to the conclusion that electrons were charged particles?

2. Describe the contribution of J.J. Thomson to atomic theory.

3. Draw a picture of the “plum-pudding” model of the atom proposed by J.J. Thomson.

4.. Draw a picture and describe the results of the “gold-foil” experiment employed by Ernest Rutherford. Draw a picture of the atom after Rutherford’s discovery of the nucleus.

The proton was determined in 1920 to have a charge equal in magnitude but opposite in sign (+) by Rutherford who noticed that ALL elements contain hydrogen-like nuclei (protons). Later, in 1932 James Chadwick was credited with discovery of the neutron.

Fill out the following table summarizing the model of the atom as it stood in 1932.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Properties of Subatomic Particles | | | | |
| Particle | Symbol | Relative Charge | Relative Mass (amu) | Actual Mass (g) |
| electron | e- |  | 1/1840 |  |
| proton |  | +1 |  | 1.673 x 10-24 |
| neutron |  |  |  |  |

**TLDR 3.3 Counting Atoms (pp. 79-90)**

The atoms of all of the elements in our universe differ from one another in the number of protons that are contained within the nucleus. It is the number of protons that fully identify the physical and chemical properties of each element. Within the population of atoms that make up an element there are differences in the number of neutrons in the nucleus. Atoms with the same number of protons (same element), but with different numbers of neutrons are called isotopes. Atoms in their “elemental state” have equal numbers of electrons and protons, which result in a charge of zero. When atoms gain electrons (anions) or lose electrons (cations) they become electrically charged and the resulting species are called ions. On the periodic table there are two types of values displayed for each element. On top of each elemental symbol is an integer value representing the atomic number of that element, this number is a counting number and therefore, has an infinite number of significant digits. The number below the symbol is called the “average atomic mass” (aka “atomic mass”) and is a weighted average of all the different isotopes that make up any sample of a particular element. Since the atomic mass value is measured, and all measurements contain some error, these values carry a finite number of significant digits.

1. Give definitions of the following terms;

***atomic number, isotope, mass number, unified atomic mass unit, average atomic mass, Avogadro’s number, mole, molar mass***

2. The isotope symbol notation includes chemical1. symbol, atomic number and mass number as shown below for the isotope Cl-36



In general, this notation, for the generic element X looks like this;



Write isotope notation for the following isotopes

a. Carbon-14 b. Uranium-235 c. Hydrogen-3 d. Boron-11

3. Complete the following table

|  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- |
| Element | Atomic # | Mass # | #p+ | #n0 | #e- | Ion? (y/n) | Charge | Isotope symbol |
|  | 1 |  |  | 2 | 1 |  |  |  |
|  |  | 36 | 17 |  |  |  | -1 |  |
| O |  |  |  | 7 |  | N |  |  |
|  | 12 | 24 |  |  | 10 |  |  |  |
|  |  |  | 9 | 10 |  |  | 0 |  |

4. Calculating atomic mass. The average atomic mass of an element is a weighted average of the masses of individual isotopes, the weighting is based on the relative abundance of each element. The relative abundance of a particular isotope is the percentage of all of the naturally occurring atoms of that element that has a particular number of neutrons.

To calculate the average atomic mass of an element that is made up of two major isotopes A and B, the following method should be employed;

Example: Given the data in the table, calculate the atomic mass of unknown Element X.

|  |  |  |
| --- | --- | --- |
| Isotope Abundance for Element X | | |
| Isotope | Mass (amu) | Percent Abundance |
| 6X | 6.015 | 7.59% |
| 7X | 7.016 | 92.41% |

Solution:

Note: The raw percentages given in the table must first be divided by a factor of 100 before performing any further math operations. Also note, we are not adding two values and then dividing by two; which would be perfectly fine if the relative abundances were 50% each.

1. Three magnesium isotopes have atomic masses and relative abundances of 23.985 amu (78.99%), 24.986 amu (10.00%), and 25.982 amu (11.01%). Calculate the atomic mass of magnesium.

**Mass-mole, mole-particle conversions**

The mole is a counting number similar to “dozen”, “score”, “gross” etc. Whereas a dozen was determined to be a convenient number useful for the counting of eggs, donuts, bagels, cookies and such, the mole is an arbitrarily chosen number that is useful for counting very, very small things like atoms and molecules. A mole is defined as the number of atoms contained in a 12.00 g sample of Carbon-12 and has a value of 6.022 x 1023.

Preparing chemical reactions in the laboratory is closely analogous to baking cookies using a recipe. Chemical equations include an “ingredient list” (the reactants) on the left side and lists the results of combination of those ingredients on the right side (the products) of the chemical equation. Although we won’t focus on the details, the conditions that must be established for a particular chemical reaction to occur are commonly listed, in short hand notation above the arrow separating reactant and product sides of the chemical equation.

**Generalized form of a chemical reaction**

(lower case letters are whole number coefficients, capital letters are chemical formulas.

**aA + bB 🡪 cC+ dD**

Example (combustion of methane CH4)

**CH4 (g) + 2 O2 (g) 🡪 CO2 (g) + H2O (g)**

*In this case CH4  and O2 are the reactants, CO2 and H2O are the products*

**Mole to mass and mass to mole conversions.**

In the example above we can interpret the balanced chemical equation as follows

*1 molecule of CH4 reacts with 2 molecules of O2 to produce 1 molecule of CO2 and 1 molecule of H2O.*

*Or*

*1 mole of CH4 molecules react 2 moles of O2 molecules to produce 1 mole of CO2 molecules and 1 mole of H2O molecules.*

This ratio of 1:2:1:1 is called the stoichiometric ratio for this reaction. In the chemistry lab, we don’t have the time to individually count atoms or molecules even though the chemical reactions we perform rely on fixed ratios of the numbers of these particles. It is far more convenient for us to weigh out a solid sample and use molar masses to convert to moles for use in interpreting chemical reactions. For liquid samples we can also relate mass to a number of moles, but we need density to do so. For gaseous species, we will soon learn that knowing pressure and volume of a substance gives us a means to calculate number of moles.

**Example:**

When aqueous (dissolved in water) solutions of NaCl and AgNO3 are mixed together, a white precipitate (cloudy solid) forms that can be filtered out, dried and then weighed on an electronic balance. The balanced chemical reaction describing this process is as follows:

**AgNO3(aq) + NaCl (aq) 🡪 AgCl(s) + NaNO3(aq)**

1. Calculate the molar mass of all of the reactants and products in this reaction.
2. A student performed this reaction and ended up with 3.46 g of AgCl. How many moles of AgCl did they make? How many formula units of AgCl does this represent?

**Practice problems**

1. Calculate the number of moles contained in a 5.00 g sample of the following substances.
   1. O2
   2. CH4
   3. H2O
   4. CO2
   5. C6H12O6 (glucose)
2. Calculate the number of molecules contained in a 5.00 g sample of the following substances
   1. Cl2
   2. H2O2
   3. HCl
   4. CaCO3
   5. NaCl
3. Calculate the mass (in g) of the following samples
   1. 3.75 x 1022 molecules of H2O
   2. 3.50 mol of CH4
   3. 0.375 mol of CO2
   4. 1.5 x 1024 molecules of H2O2
   5. 2.50 mol of NaCl