Chem400

General Chemistry

Instructor: Prof. Maddox

Note:

- Students will not be added to the class during lecture
- You must wait for your specified lab session to inquire about adding the class
- Wait-list students can obtain class materials at;

http://www.arc.losrios.edu/Faculty_Web_Pages/Michael_Maddox.htm

Introduction

Welcome to Chem 400!

- This is a very challenging class for most students
- The main goal of this class is to prepare you for the rest of the Chemistry Series, starting with Chem 401
- You will be pushed hard and expected to put in a lot of time and effort (there are **lots** of quizzes and exams)
- If you pass this class, you will be **very well prepared** for Chem 401
- However, many students fail this class…

Introduction

- What causes students to fail Chem 400?

a) Students are unprepared at the start of the semester
⇒ work through "Chem 400 Basics" worksheet on Canvas (Canvas.losrios.edu) – and review before every chapter

b) Students fail to attend or be attentive in every class

c) Students don't spend enough **quality** time reviewing their notes, working through a textbook, working on their lab reports, doing homework, and preparing for tests and exams
⇒ allow at least 1 hr to prepare for a lab, 1 – 2 hrs to complete a lab report, 1– 4 hrs per chapter for homework, about 1 hr after each class to review your notes (use chapter worksheets if necessary), at least 2 hrs/day for 7 days before a mid-term, and at least 2 hrs/day for 2 weeks before the final exam

d) Students allow too much time to pass before addressing problems
⇒ if you don't understand something we cover in class: review your notes, read a textbook, ask for help (in that order)

Class Syllabus

- Read the entire class syllabus very carefully, and talk to me about any issues right away
- There is a class schedule on the last page that shows:
  ⇒ what topic will be covered in each lecture
  ⇒ what experiments will be done in each lab
  ⇒ The date of every chapter quiz (7), mid-term exam (3), and the final exam
- A "lite" version of the syllabus is printed on the back of the schedule (p.7 of syllabus) – you will take a short test on this at the start of your second lab period.

4

General Chemistry Textbook

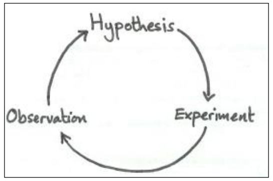
- This class has no specific official textbook, but…
  - ⇒ you are **required** to have access to a general chemistry textbook to supplement the material in the lectures
- The Openstax.org online chemistry textbook is recommended and will be used in Chem 401 classes at ARC
  - ⇒ you can download a free pdf copy
  - ⇒ you can buy a hard copy for $55 from Amazon
- Alternatively, you could buy any first year college chemistry textbook
  - ⇒ any edition would be okay
  - ⇒ older editions are significantly cheaper

Chapter 1

Getting Started

1.1 The Scientific Method

- Think of the ***scientific method*** as a cyclic process;



- The cycle usually begins with an observation, and ends when several cycles have left the hypothesis unchanged – it is then considered a theory
- Note that 'conclusion' is not really part of the scientific method. The conclusion is typically just a restatement of the most recent hypothesis

1.1 The Scientific Method

Scientific Method

- Observation – data
- Hypothesis – an explanation of the data, from which predictions can be made (should be falsifiable)
- Experiment – anything that produces data (usually set up to test the hypothesis)

After many "scientific method cycles", a hypothesis should;

1. explain all the data AND
2. make predictions with reasonable accuracy

⇒ at this point it can be called a **theory** or a **law**

- Theory – a well-tested explanation of the facts/data
- Scientific Law – a theory that can be simply stated (often mathematically, for example, $E = mc^2$)

1.2 The Classification of Matter

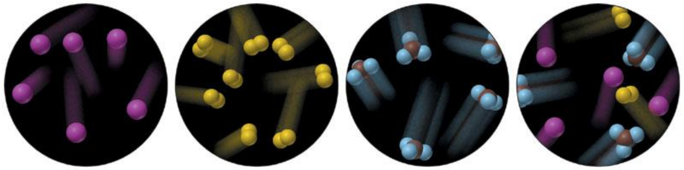
- Matter is any substance that has mass and occupies volume
- All matter is composed of atoms (or ions, which are basically just electrically charged atoms)

  ⇒ atoms are themselves composed of smaller particles
- If a 'chunk' of matter contains only one type of atom it is an element
- If the 'chunk' contains more than one type of atom or ion ***chemically bonded together***, it is a compound

  ⇒ atoms chemically bonded together form molecules
- Elements and compounds are **pure substances**

1.2 The Classification of Matter

- If the ‘chunk’ contains different elements and/or compounds, that are not chemically bonded together, it is a **mixture**



A Atoms of an element   B Molecules of an element   C Molecules of a compound   D Mixture of two elements and a compound

1.3 Measurements and Units

Precision and Accuracy: (a) Single Measurements

- A measurement is a number with a unit attached; e.g. 13.7 miles
- Measurements are more accurate if they’re closer to the true value
  - ⇒ If the true value is 39.8374 g, then;
    39.3 g is more accurate than 35.2901 g
- Measurements are more precise if they have a smaller ± range
  - ⇒ 423.52 ± 0.24 cm is more precise than 83.21 ± 0.26 cm
- Note: most measurements have an *implied* ± range, which depends on the significant figures (digits) and decimal places
  - ⇒ 130 mol is the same as 130 ± 5 mol
  - ⇒ 130. mol is the same as 130. ± 0.5 mol
  - ⇒ 130.2 mol is the same as 130.2 ± 0.05 mol

1.3 Measurements and Units

Precision and Accuracy: (b) Data Sets

- A set of data can be represented by an **average** and a **90% confidence interval** (see Stats section of lab manual for details)

  e.g. data set 1; 23.4g, 28.5g, 19.1g, 36.5g would be;

  ⇒ **26.9** ± **10.2** g

  e.g. data set 2; 20.6g, 27.3g, 21.2g, 24.8g would be;

  ⇒ **23.5** ± **4.3** g
- If the true value of the mass is 28.6 g, then **data set 1 is more accurate** because its average value (26.9 g) is closer to the true value
- **Data set 2 is more precise**, because it has a smaller ± range

1.3 Measurements and Units

Significant Figures (sig. figs. or s.f.)

⇒ 40000 has 1 s.f.	321 has 3 s.f.	0.2300 has 4 s.f.
40000. has 5 s.f.	321. has 3 s.f.	0.23 has 2 s.f
40.000 has 5 s.f	3.21 has 3 s.f.	0.0023 has 2 s.f.

Sig. Fig. Rules

1) a number with no decimal point;
⇒ Start from the left and stop counting at the last non-zero number

2) a number with a decimal point;
⇒ Start from the right and stop counting at the last non-zero number

1.3 Measurements and Units

- Scientific notation (exponential notation) utilizes the significant digits in a measurement multiplied by a power of ten.
- Only the significant digits are shown, and these are usually expressed as a number between 1 and 10.

power of 10

$$D\,.\,D\,D \times 10^{n}$$

significant digits

- Example 1: $456.1 = 4.561 \times 10^{2}$
- Example 2: $0.039 = 3.9 \times 10^{-2}$
- Example 3: $6400 = 6.4 \times 10^{3}$

1.3 Measurements and Units

Sig. Figs. and Calculations

- When performing calculations, the precision of the answer is dictated by the measurement with the ***most uncertainty*** (the weakest link in the chain)

5.15 cm × 2.3 cm = 11.845 cm^2 ⇒ 12 cm^2

2 m ÷ 0.45328 s = 4.41228… m/s (ms^{-1}) ⇒ 4 m/s

80.800 s - 3.9 s = 76.900 s ⇒ 76.9 s

5100 m + 133 m = 5233 m ⇒ 5200 m

Total mass of 7 stones = 58.32 g

What is the average mass of one stone?

58.32 g ÷ 7 = 8.33142… g ⇒ 8.331 g

1.3 Measurements and Units

- There are different rules for multiplication/division and addition/subtraction

⇒ Multiplying or dividing: the answer has the same number of sig. figs. as the number with the fewest sig. figs. in the calculation

⇒ Adding or subtracting: the final answer has the same precision (±) as the number in the calculation with the lowest precision (biggest ±) - if the numbers have decimals, the answer has the same number of decimals as the number with the least decimals in the calculation

Note: When we count something, it is an ***exact number***; *e.g. 10 oranges, 43 people, 7 stones*

When a value is defined, it is an ***exact number***; e.g. there are exactly 1000 m in 1 km

⇒ Significant digit rules do not apply to exact numbers

1.3 Measurements and Units

- An experiment produces a result of 2.929 and the expected (actual) value is 2.91652, what is the experimental % error?

$$\%\ error = \frac{(2.929 - 2.91652)}{2.91652} \times 100\%$$

- This calculation uses a combination of division/multiplication and subtraction – you must use the appropriate rule for each part to find the correct s.f. or d.p. for that part, **but only round off at the end**

2.929 – 2.91652 = 0.01248 (answer good to 3 d.p.)

0.01248 ÷ 2.91652 = 0.004279 (answer good to 2 s.f.)

0.004279 x 100% = 0.4279% (answer to 2 s.f.) 100% is an exact number

Answer has 2 s.f., so it rounds off to 0.43%

Note: rounding each step gives 0.012 ÷ 2.961652 = 0.0041 ⇒ 0.41%

1.3 Measurements and Units

Quick Note (Warning) about Run-On Calculations

- Each step in a multi-step calculation (like the one on the previous slide) should be written separately
- If you put separate steps together, as in the following example, you are committing a "math crime"

⇒ 256 – 43 = 213 x 172 = 36,636

- This is **incorrect**, because it states that 256 – 43 = 36,636 (if a = b = c, then a = c)
- You cannot do one step (256 – 43 = 213) and then simply continue the calculation – you must write the steps separately
- **If you write run-on calculations in labs, quizzes, or exams, you will lose points**

1.3 Measurements and Units

- SI units use *prefixes* to enlarge or reduce the basic units (g, s, mol, m, L, etc.)
- Some are familiar;

⇒ A kilometer (km) is 1,000 meters

1 km = 1,000 m or 0.001 km = 1 m

- Some are less familiar;

⇒A microliter (μL) is 1/1,000,000 of a liter

1 μL = 0.000 001 L or 1,000,000 μL = 1 L

1.3 Measurements and Units

- The *prefixes* you will need to know for this class are;

⇒ Mega (M)	1 Mg = 1,000,000 g
⇒ Kilo (k)	1 kg = 1,000 g
⇒ deci (d)	1 g = 10 dg
⇒ centi (c)	1 g = 100 cg
⇒ milli (m)	1 g = 1,000 mg
⇒ micro (μ)	1 g = 1,000,000 μg
⇒ nano (n)	1 g = 1,000,000,000 ng

These are called "unit equations"

1.3 Measurements and Units

- You will come across some units that are combinations of other units; for example, energy is measured in joules, J
- These are called **compound units**;

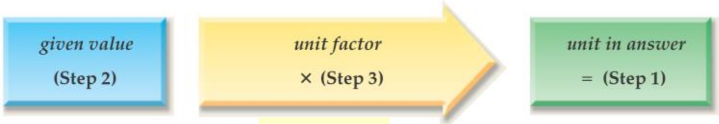
⇒ $1\ \mathrm{J} = 1\ \mathrm{kg\ m^2 / s^2}$ (which is the same as $1\ \mathrm{kg\ m^2\ s^{-2}}$)

- Compound units can be written in several different ways;

$$\frac{\mathrm{L \cdot atm}}{\mathrm{K \cdot mol}} = \mathrm{L\ atm / K / mol} = \mathrm{L\ atm\ K^{-1}\ mol^{-1}}$$

1.4 Unit Factor Calculations

- The unit factor (conversion factor) method is useful when converting units (and we will use it for much more, later on);
- Convert 78 mL into L (1 L = 1000 mL or 1 mL = 0.001 L)



78 mL $\quad \times \dfrac{1\,\text{L}}{1000\,\text{mL}} \quad$ L

$$78\ \cancel{\text{mL}} \times \frac{1\,\text{L}}{1000\,\cancel{\text{mL}}} = 0.078\ \text{L}$$

1.4 Unit Factor Calculations

- Sometimes, more than one unit factor is needed;

⇒ How many cm in 2.00 ft? (12 in = 1 ft, 2.54 cm = 1 in)

$$2.00\,\cancel{\text{ft}} \times \frac{12\,\cancel{\text{inch}}}{1\,\cancel{\text{ft}}} \times \frac{2.54\,\text{cm}}{1\,\cancel{\text{inch}}} = 61.0\,\text{cm}$$

⇒ What is 4981 cm^3 in cubic meters, m^3?

$$4981\,\cancel{\text{cm}^3} \times \frac{1\,\text{m}}{100\,\cancel{\text{cm}}} \times \frac{1\,\text{m}}{100\,\cancel{\text{cm}}} \times \frac{1\,\text{m}}{100\,\cancel{\text{cm}}} = 4.981\times10^{-3}\ \text{m}^3$$

⇒ What is 95 km/hour in meters per second (m/s)?
(1 km = 1000 m and 1 h = 3600 s)

$$\frac{95\,\cancel{\text{km}}}{\cancel{\text{hr}}} \times \frac{1000\,\text{m}}{1\,\cancel{\text{km}}} \times \frac{1\,\cancel{\text{hr}}}{3600\,\text{s}} = \frac{26\,\text{m}}{\text{s}}$$

1.4 Unit Factor Calculations

- Density (and other properties with compound units) can be used as a unit factor
- If the density of an acid is 1.84 g/mL, how many grams of acid are in 1275 mL?

$$1275\,\text{mL} \times \frac{1.84\,\text{g}}{1\,\text{mL}} = 2350\,\text{g}$$

- If the density of a rock is 5.23 g/cm^3, what is the volume of a 75.0 g rock sample?

$$75.0\,\text{g} \times \frac{1\,\text{cm}^3}{5.23\,\text{g}} = 14.3\,\text{cm}^3$$

- Note: the density itself can be calculated if the mass and volume are known;

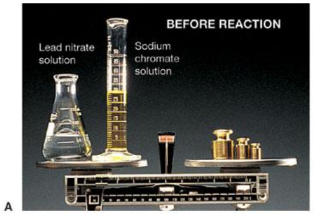
$$\text{Density} = \frac{\text{mass}}{\text{volume}}$$

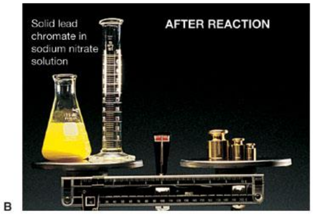
Chapter 2

Atoms and Elements

2.1 Atomic Theory: Basic Laws

- The Law of Mass Conservation
  - ⇒ ***the total mass of substances does not change during a chemical reaction***
  - ⇒ ***the mass of all products is equal to the mass of all reactants that <u>actually react</u>***





2.1 Atomic Theory: Basic Laws

- Example Law of Mass Conservation Calculation
  - ⇒ In a reaction between Zn and I_2, the expected mass of ZnI_2 produced can be found using the Law of Mass Conservation

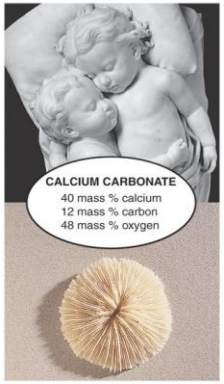
$$I_2 + Zn \rightarrow ZnI_2$$

- When 10.0 g of I_2 reacts with 7.0 g of Zn, all the I_2 is used up and 4.4 g of Zn remains unreacted
  - ⇒ 10.0 g of I_2 and 2.6 g of Zn reacted, so the product mass should be 12.6 g

<u>Note</u>: this approach relies on several accurate mass measurements – stoichiometry is generally a better way to predict product mass

2.1 Atomic Theory: Basic Laws

- The Law of Definite Proportions (Constant Composition)
  - ⇒ ***no matter what its source, a particular compound is always composed of the same elements in the same mass ratio***



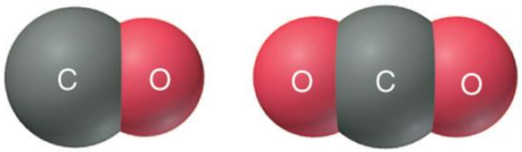
2.1 Atomic Theory: Basic Laws

- The Law of Multiple Proportions
  - ⇒ ***if elements A and B react to form two different compounds, the different masses of B that combine with a fixed mass of A can be expressed as a ratio of small whole numbers***
- Consider two compounds composed of carbon and oxygen;
  - ⇒ carbon oxide I: 1.00g carbon, 1.33g oxygen
  - ⇒ carbon oxide II: 1.00g carbon, 2.66g oxygen
  - ⇒ 1.33g : 2.66g = 1 : 2 (ratio of small whole numbers)

2.1 Atomic Theory: Basic Laws

- For the same mass of carbon, carbon oxide II contains exactly 2 times as much oxygen as carbon oxide I (not 1.342 times or 2.109 times)
- This suggests that elements come in ***chunks*** (they can't be infinitely divided)
  - ⇒ you can combine one ***chunk*** of oxygen with carbon, or two chunks of oxygen with carbon, but not half a ***chunk*** (or any other fractional amount)
  - ⇒ there must be a basic elemental unit – the atom

2.1 Atomic Theory: Basic Laws



Carbon oxide I (carbon monoxide)

Carbon oxide II (carbon dioxide)

2.2 Atomic Theory: Dalton's Model

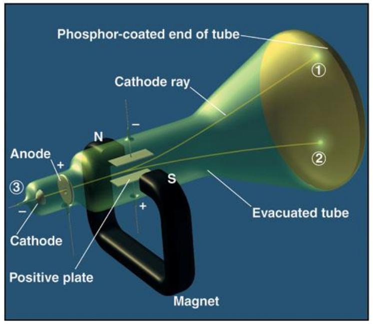
Dalton's Atomic Model

- The first comprehensive, modern theory (model) of the atom was proposed by John Dalton in the early 1800s:
  1) Elements, in their purest state, consist of indivisible particles called atoms (from Law of Multiple Proportions)
  2) The atoms of a specific element are all identical
  3) Atoms of different elements can be distinguished by their atomic masses
  4) Atoms of elements combine to form chemical compounds in Definite Proportions
  5) Atoms can neither be created nor destroyed in a chemical reaction, only their grouping changes (Law of Cons. Of Mass)

     e.g. C in CH_4 becomes C in CO_2 when you burn CH_4 in air

2.3 Atomic Theory: Thomson's Model

- About 100 years later, J.J. Thomson used a cathode ray tube…



Tiny, negatively-charged particles discovered: the electron

⇒ Much lighter than a hydrogen atom

⇒ Atoms are NOT indivisible

OBSERVATION	HYPOTHESIS
1. Ray bends in magnetic field	Consists of charged particles
2. Ray bends toward positive plate in electric field	Consists of negative particles
3. Ray is identical for any cathode	Particles found in all matter

2.3 Atomic Theory: Thomson's Model

- The charge of an electron was calculated from **Millikan's oil-drop experiment** (see Openstax, p.74)
  - ⇒ oil drops became negatively charged due to whole numbers of electrons sticking to them (for example, there could be 5 electrons, or 237, but not 3.4)
  - ⇒ the highest common factor of the charges on all the different oil drops was -1.602 x 10^{-19} C
  - ⇒ this is the charge of a single electron
- These new observations allowed J. J. Thomson to correct Dalton's idea that all atoms are indivisible
  - ⇒ Dalton's other ideas were unchanged

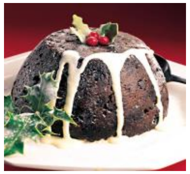
2.3 Atomic Theory: Thomson's Model

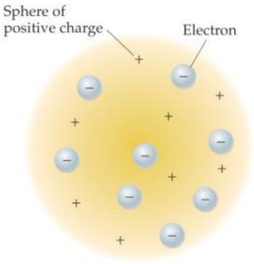
The "Plum Pudding" Atomic Model

1) Atoms are made of tiny, negatively charged particles (electrons), evenly distributed throughout a **homogeneous** sphere of positive charge

Note: there are no positive particles, or any nucleus in this model

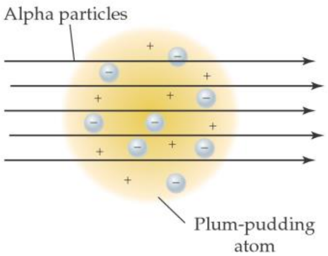
- This atomic model got its common name because it reminded people of a plum pudding





2.4 Atomic Theory: Rutherford's Model

- According to the "Plum Pudding Model", big, heavy alpha particles ($^4_2He^{2+}$) should blast right through the "Plum Pudding" cloud.
  - ⇒ like bullets through a plum pudding



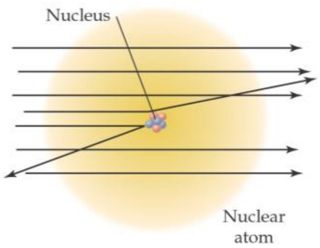
- Rutherford experiment (1910)

2.4 Atomic Theory: Rutherford's Model

BUT…

- **Most** of the alpha particles passed through the atom because an atom is mostly empty space
- The alpha particles that bounced backwards did so after striking something very small and very heavy
- At the center of an atom is the ***atomic nucleus*** which contains all the atom's protons

⇒ protons are much heavier than electrons



2.4 Atomic Theory: Rutherford's Model

- Based on the heaviness of the nucleus, it had to contain other particles in addition to protons
- Neutrons, n^0 (or just n), are also in the nucleus

⇒ A neutron is about the same mass as a proton, but without any charge

Summary of Subatomic Particles

Subatomic Particle	Mass (amu)	Relative Charge	Location
Electron	~1/1800	-1	Around nucleus
Proton	~1	+1	Nucleus
Neutron	~1	0	Nucleus

Note: (1g = 6.02 x 10^{23} amu)

2.4 Atomic Theory: Rutherford's Model

Rutherford's Atomic Model

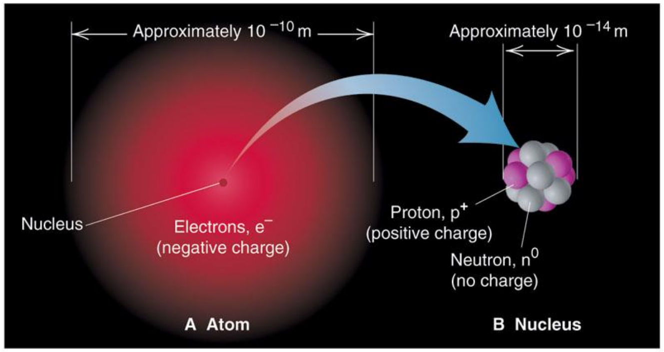
1) Atoms are mostly empty space containing rapidly moving, negatively charged electrons
2) At the center of an atom there is a tiny cluster of neutrons and positively charged protons, called the nucleus

⇒ the diameter of the nucleus is about 10,000 times smaller than the diameter of the atom

3) Atoms are electrically neutral

⇒ negatively charged electrons are distributed around a positively charged nucleus

2.4 Atomic Theory: Rutherford's Model



2.5 Subatomic Particles

- Atomic number (Z) is the number of protons in an atom

  ⇒ all atoms of the same element have the same atomic number

- Mass number (A) is the number of protons + neutrons in an atom
- We use ***<u>atomic notation</u>*** to show the <u>atomic number</u> and <u>mass number</u> of an atom:

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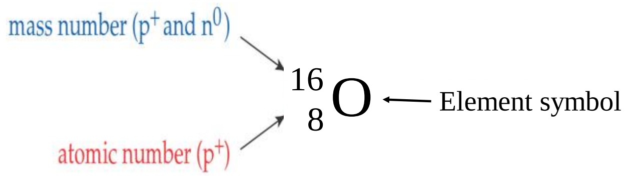


2.5 Subatomic Particles

- Unlike the number of protons, the number of neutrons is <u>not fixed</u> for each element.

  ⇒ carbon atoms can have 6, 7, or 8 neutrons

  ⇒ carbon atoms always have 6 protons, so the **mass number** can be 12, 13, or 14 (depending on the number of neutrons).

2.5 Subatomic Particles

- Atoms of an element that have <u>different</u> **mass numbers** are called **<u>isotopes</u>**.

  ⇒ carbon has 3 naturally occurring isotopes; carbon-12, carbon-13, and carbon-14

- Different **Isotopes** of the same element are <u>chemically identical</u> (but not physically identical)

  ⇒ 'heavy water' contains a heavier-than-normal isotope of hydrogen (deuterium)

  ⇒ in <u>chemical reactions</u>, it behaves exactly the same as regular water, but it has a **higher density**

2.6 Atomic Mass

- Atomic mass of an element is the <u>weighted average mass</u> of all the stable isotopes of that element

  ⇒ this is <u>not the same</u> as the mass number, which refers to the number of <u>nucleons</u> (protons + neutrons) in a single atom of a specific isotope

  Example: magnesium (Mg) has <u>three</u> stable isotopes

  ^{24}Mg (23.985 amu, 78.99% abundance)

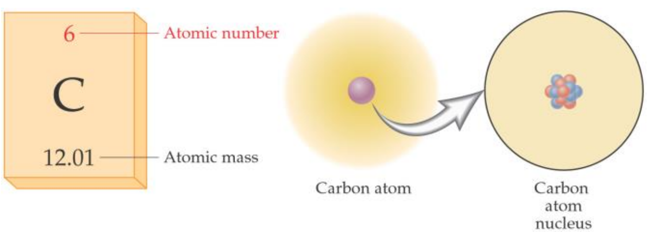
  ^{25}Mg (24.986 amu, 10.00% abundance)

  ^{26}Mg (25.983 amu, 11.01% abundance)

  Atomic mass = (^{24}Mg fractional abundance x ^{24}Mg mass) +
  (^{25}Mg fractional abundance x ^{25}Mg mass) +
  (^{26}Mg fractional abundance x ^{26}Mg mass)

2.6 Atomic Mass

- The periodic table shows the atomic number, symbol, and atomic mass for each element.



2.6 Atomic Mass

- Atomic mass = (0.7899 x 23.985 amu) +
  (0.1000 x 24.986 amu) +
  (0.1101 x 25.983 amu) =
  <u>24.31 amu</u> (see periodic table)

Note: unlike simple (non-weighted) averages, you <u>do not</u> divide by anything in the <u>weighted average</u> calculation