Announcements for Thursday, 12SEP2024

For those who just recently joined the class:

- Check Canvas Announcements and e-mails often and read through all the posted material as soon as possible to get current with the class

For everyone:

- Homework Assignments available on Canvas/eLearning
  - Week 2: Graded and Timed Quiz 2 – "Essentials" due **Monday, 16SEP2024, at 6:00 PM (EDT)**
  - Week 2: **Readiness Assessment** will be **re-opened** on **tomorrow for 24 hours**
  - Week 2: *Study Skills* and *Time Management* Digital Badge Assignments due **tomorrow at 11:59 PM (EDT)**
  - Week 3: Beginning of Semester Chemistry surveys due **Monday, 16SEP2024, at 11:59 PM (EDT)**
  - Week 3: *Metacognition* Digital Badge Assignment due **Friday, 20SEP2024, at 11:59 PM (EDT)**
- In-person/online recitations begin this week
  - Students interested in ALWs should attend regular recitations until officially accepted
- First Day Course Materials – See Canvas announcement about opting-out (deadline: 17SEP2024)

ANY GENERAL QUESTIONS? Feel free to see me after class!

Isotopes

Dalton: "All atoms of a given element have the same mass and other properties…"

…Not Quite

- **isotopes =** atoms with the same number of protons but different number of neutrons (and, therefore, slightly different masses)
  - different isotopes of an element generally exhibit the same ***chemical*** behavior
- a naturally occurring sample of a given element may be made up of more than one isotope
  - example: carbon (6 p^+) with 6, 7, or 8 neutrons
- **mass number (A) =** number of protons + number of neutrons
- **natural abundance =** the relative amount of an isotope in a naturally occurring sample of an element; usually given as percentages

- symbolizing isotopes: ${}^{A}_{Z}X$notation vs. X-A notation

Try This On Your Own

Use a periodic table and A/Z notation to give the symbol for the following species or fill-in the missing information

number of protons	number of neutrons	number of electrons	symbol
9	10	9	
2	2	0	
33	42	36	
			${}^{109}_{52}\mathrm{X}^{2-}$
			${}^{55}_{?}Mn^{3+}$

Atomic Mass

- the atomic mass of each element is listed directly beneath its symbol on the periodic table
    - NOTE THE LACK OF UNITS! More to come about this…

- it is a WEIGHTED AVERAGE (?!?) of the masses of the isotopes that compose that element
    - the masses of isotopes with greater natural abundances impact the average atomic mass more so than isotopes with lesser natural abundances
    - a common example of a weighted average: your GPA

- example: carbon (atomic mass = 12.01 amu)

6
C
12.01
carbon

Calculating Average Atomic Mass of Carbon

isotope	mass (amu)	% abundance	fractional abundance (F.A.)
carbon-12	12.0000	98.93%	0.9893
carbon-13	13.0034	1.07%	0.0107

$$\begin{array}{c}\text{Average} \\ \text{Atomic Mass}\end{array} = [(\mathrm{F.A.}\ {}^{12}_{6}\mathrm{C})(\mathrm{mass}\ {}^{12}_{6}\mathrm{C})] + [(\mathrm{F.A.}\ {}^{13}_{6}\mathrm{C})\ (\mathrm{mass}\ {}^{13}_{6}\mathrm{C})]$$

$$= [(0.9893)(12.0000)] + [(0.0107)\ (13.0034)]$$

$$= [11.87] + [0.139]$$

$$= 12.01\ \mathrm{amu}$$

Calculating the Average Atomic Mass of an Element in General

definitely an important skill for this course!

isotope	mass (amu)	abundance
isotope-1	mass1	%1
isotope-2	mass2	%2
isotope-3	mass3	%3

$$\begin{matrix}\text{Average}\\ \text{Atomic Mass}\end{matrix} = \underbrace{\left[\left(\frac{\%1}{100}\right)(\text{mass1})\right]}_{\text{isotope-1}} + \underbrace{\left[\left(\frac{\%2}{100}\right)(\text{mass2})\right]}_{\text{isotope-2}} + \underbrace{\left[\left(\frac{\%3}{100}\right)(\text{mass3})\right]}_{\text{isotope-3}} \ldots$$

... etc

Remember: all percent abundances must add to **100%**

Try This On Your Own

Magnesium has three naturally occurring isotopes:

magnesium-24 (23.99 amu, 78.99% abundant)

magnesium-25 (24.99 amu, 10.00% abundant)

magnesium-26 (25.98 amu)

Calculate the average atomic mass of magnesium and compare it to the value given on your periodic table.

Beware of a Common Mistake!

Atomic Number (Z) vs. Atomic Mass vs. Mass Number (A)
Don’t confuse them!

the Mole (mol)

- it is a unit of amount (i.e., a unit of “how many?”…a countable amount)
- note that “mol” is the abbreviation of “mole”, NOT “molecule”
- you may be already familiar with some common units of amount
  - pair, dozen, baker’s dozen, gross
- conceptually we must distinguish between an actual number of entities given as a whole number and a number of entities given as an amount
  - 12 eggs vs 1 dozen eggs (BE CARFEUL OF THE DETAILS!!)
- **the mole is simply a unit of amount** analogous to the dozen, except that the number associated with the mole is much larger than 12

the Mole (continued)

$$1 \text{ mol} = 6.022\ 140\ 76\times10^{23}$$ **entities (?!?)**

- can be atoms, ions, molecules, people, books…whatever!

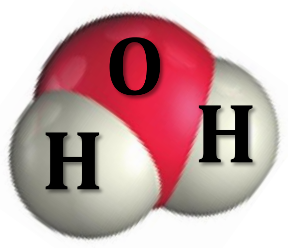
602 214 076 000 000 000 000 000.

- $6.022\ 140\ 76\times10^{23}$ is called Avogadro's number (N_A).
  - Like 12, N_A is ultimately a (HUGE) whole number
  - the magnitude of N_A…
- if you can work with the dozen, you can work with the mole
  - just like the dozen, you can have fractions of a mole to represent a whole number of things (0.5 dozen eggs is still 6 whole eggs)
- Be Careful of what you have a mole of and what you have N_A of …
  - 1 mole of gold vs. 1 mole of CO_2 vs. 1 mole sodium chloride (NaCl)
- N_A can be used as a ***conversion factor***
  - converts number of entities ↔ moles of entities

the Mole (continued)

1 mol = 6.022×10^{23} entities

consider 1 billion (1×10^9) molecules of water



determine the following

- the amount of water in **dozens** $1 \times 10^9\ \mathrm{H_2O\ molecules} \times \dfrac{1\ \mathrm{dozen}}{12\ \mathrm{molecules}} = \mathbf{8.33 \times 10^7\ dozen\ molecules}$
- the amount of water in **moles** $1 \times 10^9\ \mathrm{H_2O\ molecules} \times \dfrac{1\ \mathrm{mole\ water}}{6.022 \times 10^{23}\ \mathrm{molecules}} = \mathbf{1.66 \times 10^{-15}\ moles\ of\ water}$
- the **number** of hydrogen atoms $1 \times 10^9\ \mathrm{H_2O\ molecules} \times \dfrac{2\ \mathrm{H\ atoms}}{1\ \mathrm{molecule\ H_2O}} = \mathbf{2 \times 10^9\ H\ atoms}$
- the number of **moles** of hydrogen $2 \times 10^9\ \mathrm{H\ atoms} \times \dfrac{1\ \mathrm{mole\ hydrogen}}{6.022 \times 10^{23}\ \mathrm{H\ atoms}} = \mathbf{3.32 \times 10^{-15}\ moles\ of\ H}$
- the total **number** of atoms $1 \times 10^9\ \mathrm{H_2O\ molecules} \times \dfrac{3\ \mathrm{atoms}}{1\ \mathrm{molecule\ H_2O}} = \mathbf{3 \times 10^9\ atoms}$
- the number of **moles** of atoms $3 \times 10^9\ \mathrm{atoms} \times \dfrac{1\ \mathrm{mole\ atoms}}{6.022 \times 10^{23}\ \mathrm{atoms}} = \mathbf{4.98 \times 10^{-15}\ moles\ of\ atoms}$

DETAILS MATTER!!

Molar Mass ($\mathcal{M}$) of an Element

6
C
12.01
carbon

$$\mathcal{M} = \frac{mass\ substance\ (g)}{amount\ substance\ (mol)} = {}^{g}/_{mol}$$

- since atoms are so small, we are unable to count them out. The only way we can "count" out a number of atoms is by weighing them
  - a substance's MOLAR MASS is mass of a substance required to provide N_A of that substance
  - it is numerically equal to the atomic mass of the substance
  - BE CAREFUL OF UNITS NOW: 12.01 **amu/atom** vs. 12.01 **g/mol**
- can be used as a ***conversion factor***
  - converts mass of a substance $\leftrightarrow$ moles of a substance

The Conversion Factor Approach to Problem Solving (continued)

- What volume, in mm^3, is occupied by 9.55×10^{15} aluminum atoms given that density of Al = 2.70 g/cm^3 and the molar mass of Al = 26.98 g/mol ?

$$9.55\text{x}10^{15}\ \cancel{\text{Al atoms}} \times \frac{1\ \cancel{\text{mole Al}}}{6.022\text{x}10^{23}\ \cancel{\text{Al atoms}}} \times \frac{26.98\ \cancel{\text{g}}}{1\ \cancel{\text{mole Al}}} \times \frac{1\ \cancel{\text{cm}^3}}{2.70\ \cancel{\text{g}}} \times \frac{1000\ \text{mm}^3}{1\ \cancel{\text{cm}^3}} = 1.58\text{x}10^{-4}\ \text{mm}^3$$

More Conversion Practice Problems

- Convert 22.5 km^3 to ft^3 (2.54 cm = 1 in)
- A sample of uranium contains 1.4×10^{20} atoms. How many moles of uranium is this?
- How many dozens of silver atoms are in 0.214 moles of silver?
- Calculate the mass, in mg, of 2.25×10^{26} magnesium atoms.
- A drop of mercury has a volume of 22.0 μL and a density of 13.55 g/cm^3. How many atoms of mercury are contained within this drop?
- A 1.550-m^3 sample of a pure metal having a density of 21.40 g/cm^3 is known to contain 1.024×10^{29} atoms. With this data and a periodic table, identify the metal.