The Mole

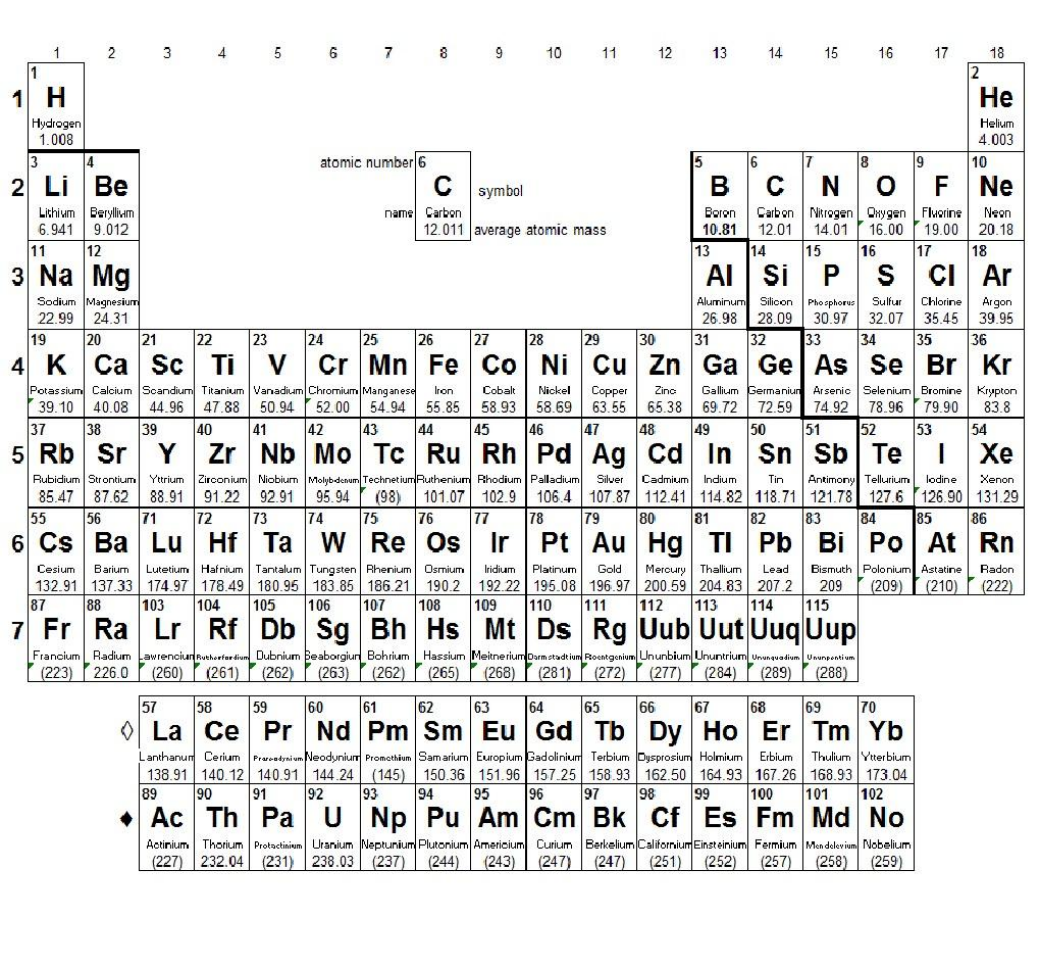
How do we connect atoms and molecules to something we can actually measure?

Counting by mass

- How did we count the number of beans in the jar?
- Similar method for counting atoms
- Atoms combine in specific numeric ratios
  - Too small and too many to actually count
- We must count them by using their average mass – just like the beans

Average Atomic Mass

- Where can we find information about the average mass of atoms?



Average Atomic Mass

- Periodic table gives average mass of an atom
  - Weighted by percent of each stable isotope
- Average atomic masses are given in AMU
- Atomic Mass Unit (AMU) = 1/12 the mass of a Carbon-12 atom or 1.6605×10^{-27} kg
- What's the problem with using AMU?
- We need a way to ramp up the numbers to something we can actually measure.

What do they have in common?

- What do the items I handed out have in common with each other?
- 1 dozen = 12 items
- 1 mole = 6.022×10^{23} items
  - Avagadro's number
- What is special about this number?
- Allows us to use the same numbers on the periodic table for single atoms or moles of atoms
  - Ex, 1 carbon atom → average mass = 12.01 AMU

    1 mole of carbon atoms → average mass = 12.01 grams

The Mole

- 1 atom H = 1.01 AMU
- 10 atoms H = 10.10 AMU
- 1000 atoms H = 1010 AMU
- 1,000,000 atoms H = 1010000 AMU
- 6.022×10^{23} atoms = 1.01 grams
- 6.022×10^{23} = Avagadro's Number = 1 mole
- This number allows us to use the average atomic masses from the PT with a unit that we can actually measure (grams)

Equivalence Statements

- 1 mole = 6.022×10^{23} particles
  - Particles = atoms, molecules, formula units
- 1 mole = 1 molar mass
  - Molar mass depends on formula
    - Elements → average atomic mass from periodic table
    - Compounds → sum of all elements in formula (x subscript)

Conversion Factors

- Using these equivalence statements, we can make conversion factors
- 1 mole = 6.022×10^{23} particles

$$\frac{1 \text{ mole}}{6.022 \times 10^{23} \text{ particles}} \qquad \frac{6.022 \times 10^{23} \text{ particles}}{1 \text{ mole}}$$

- 1 mole = 1 molar mass

$$\frac{1 \text{ mole}}{\text{molar mass}} \qquad \frac{\text{molar mass}}{1 \text{ mole}}$$

Mole Conversions

- Converting between moles, grams, and # of particles can be done using these conversion factors
  - Same method as converting units
- How many atoms are there in 2.54 moles of Aluminum?

$$\frac{2.54 \text{ moles Al}}{} \left| \frac{6.022 \times 10^{23} \text{ atoms Al}}{1 \text{ mole Al}} \right|$$

- What is the mass of 2.54 moles of Aluminum?

$$\frac{2.54 \text{ moles Al}}{} \left| \frac{26.98 \text{ g Al}}{1 \text{ mole Al}} \right|$$

Mole Conversions

- How would the conversion problems change if I substitute sodium for aluminum?

Mole Conversions

- 17.95 grams of potassium oxide is how many moles?

- If you have 3.825×10^{24} molecules of carbon dioxide, how many moles do you have?

Mole Conversions

- What is the mass of 4.112×10^{22} atoms of zinc?
- How many molecules of dinitrogen pentoxide are there in 1.074 grams?