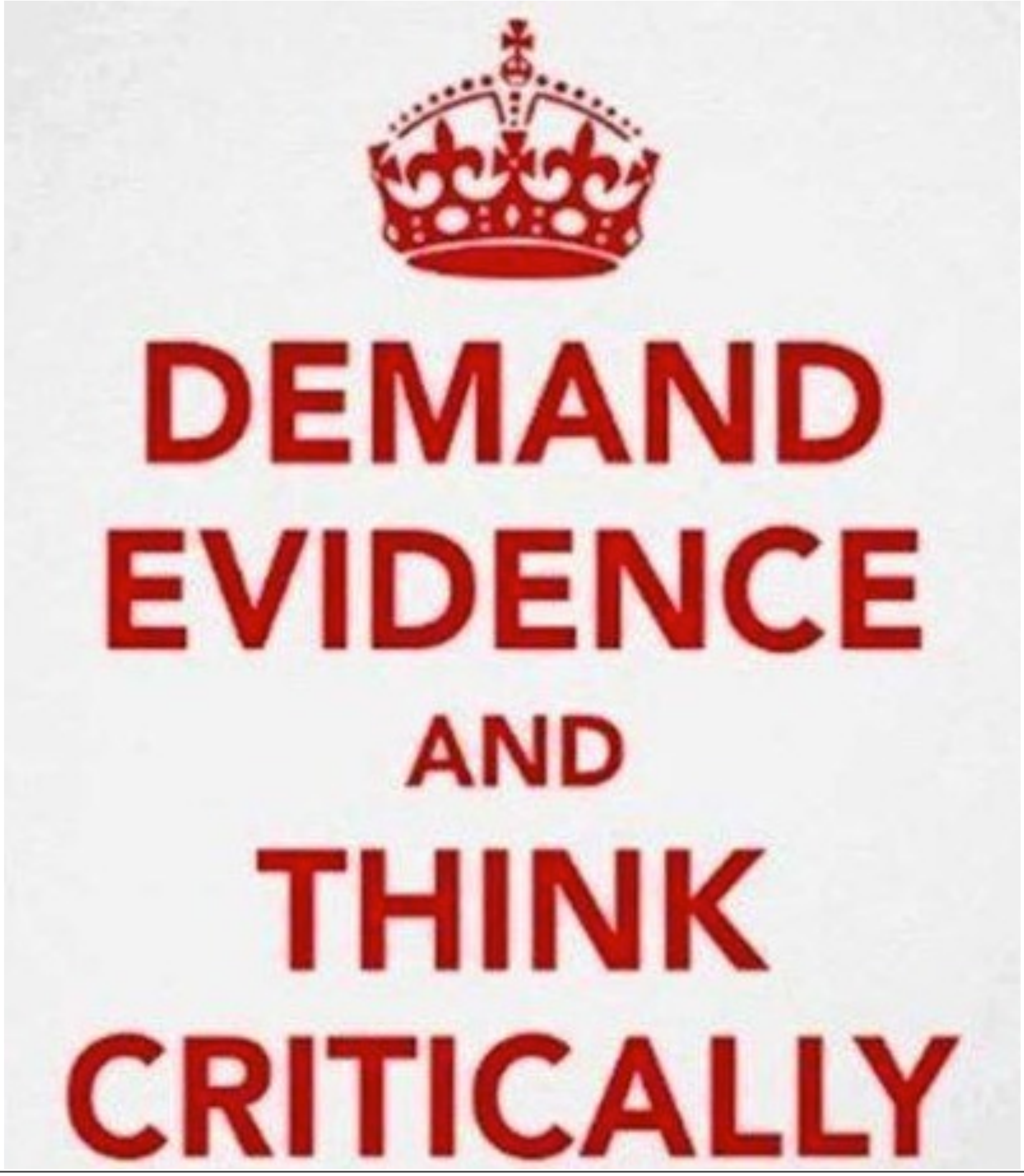
Do Now: In your notes, write down what this has to do with reasoning.



Ch. 10 Homework

- Homework Due Wednesday: Ch problems #s 50-54, 58, 59, 62-66, 68, 69.

Ch. 10: Chemical Quantities

- The Mole: Avogadro's Number.
- Mole Mass and Volume Relationships.
- Percent Composition and Chemical Formulas.

Measuring Matter

- 3 main methods for measuring matter:
  - Count it (#)
  - Mass (Kg)
  - Volume (_m^3 or _L)

Relating Quantities of Measurement

- We can convert one measurement of matter to another.
- Ex: number to weight
- This is an estimation based on averages.

Dimensional Analysis

Starting amount	Equal amounts	End Amount
24 inches	1 foot	= feet
	12 inches	

Starting amount	Equal amounts	End Amount
24 ~~inches~~	1 foot	= **2** feet
	12 ~~inches~~	

Examples

- How many eggs are in three dozen?
- How many cm in 1 km?
- How many ml in 500 L?
- How many cycles in 2.6GHz?

How many apples in a bushel?

- 1 dozen apples = 12 apples
- 1 dozen apples = 2.0 Kg of apples
- 1 dozen apples = 0.2 bushels of apples
- Use Dimensional Analysis (DA) to solve for the number of apples in a bushel as well as the weight of a bushel of apples.

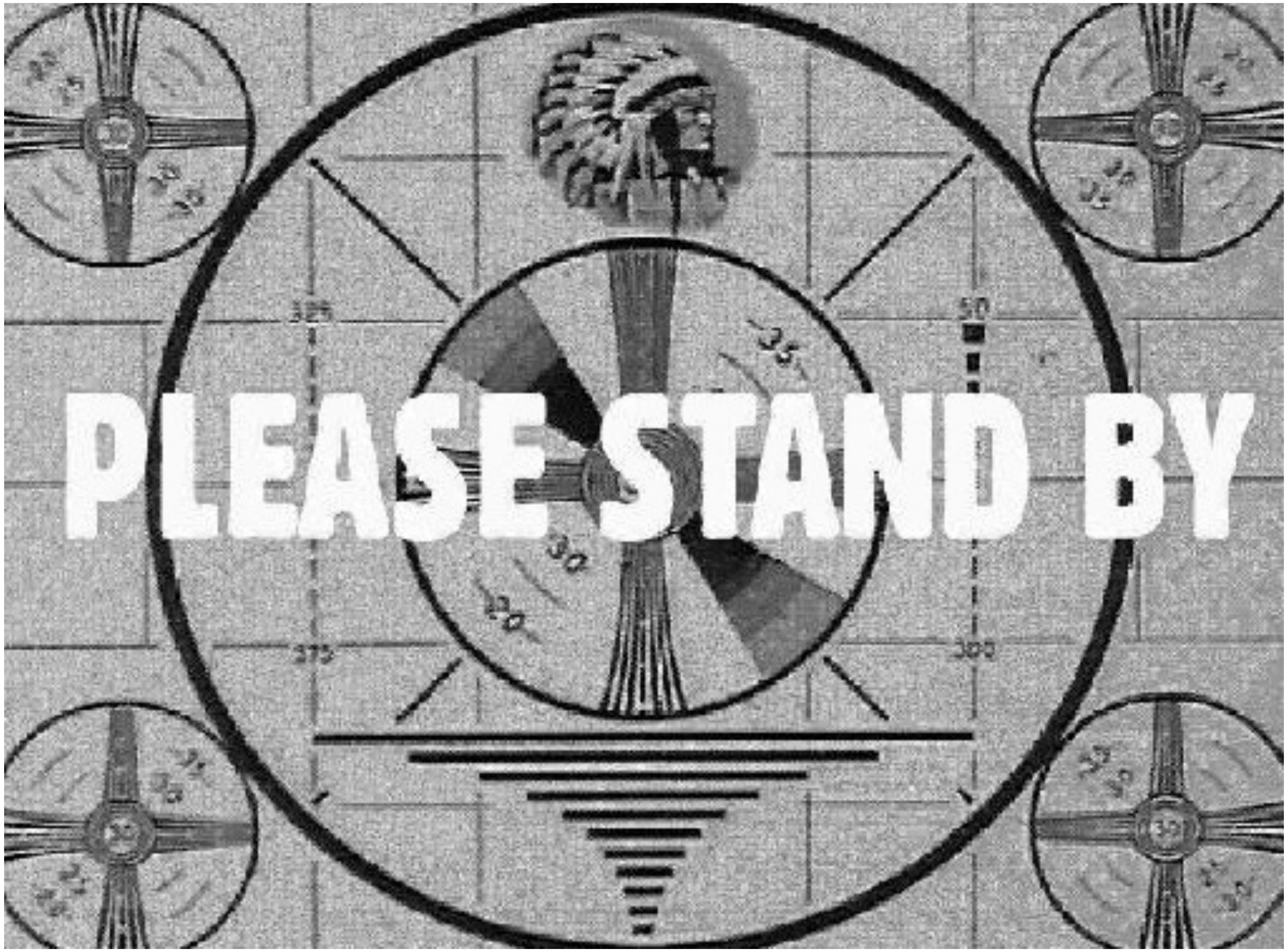
Estimation

- Use dimensional analysis to estimate the number of pizzas needed for a party given the following data:
  - There are 5 families coming to the party.
  - Each family has about 4 people.
  - The average person eats about 2.5 slices of pizza.
  - Each pizza has eight slices.

The Mole

The Mole

- Relates Weight, Atomic Mass and the Number of particles of a substance.
- 1 Mole = 1 gram of AMU
- If the AMU of Oxygen is 16, then a sample of 16g of oxygen contains 6.02×10^{23} atoms of Oxygen.
- If you know the atomic mass of a substance, then you can calculate the # of representative particles you have.



Intermission

Molar Mass

- The weight of one mole of an element.
- Finding the molar mass of a compound:
  - Find the Atomic Mass of the compound.
  - Take the formula of the compound that you are working with and figure out how many of each particle are in the compound.
  - Find the atomic mass of each element and total up the mass.

Mass of one Mole

- The mass of one mole of a particle (atom, ion, molecule, subatomic particle) is equal to the atomic mass of the particle. Instead of AMU it's in grams.
- CO^2: Carbon Dioxide. AMU of the compound is about 40 AMU.
- That means that one mole of CO^2 weighs 40g.

Mole-Mass Relationship

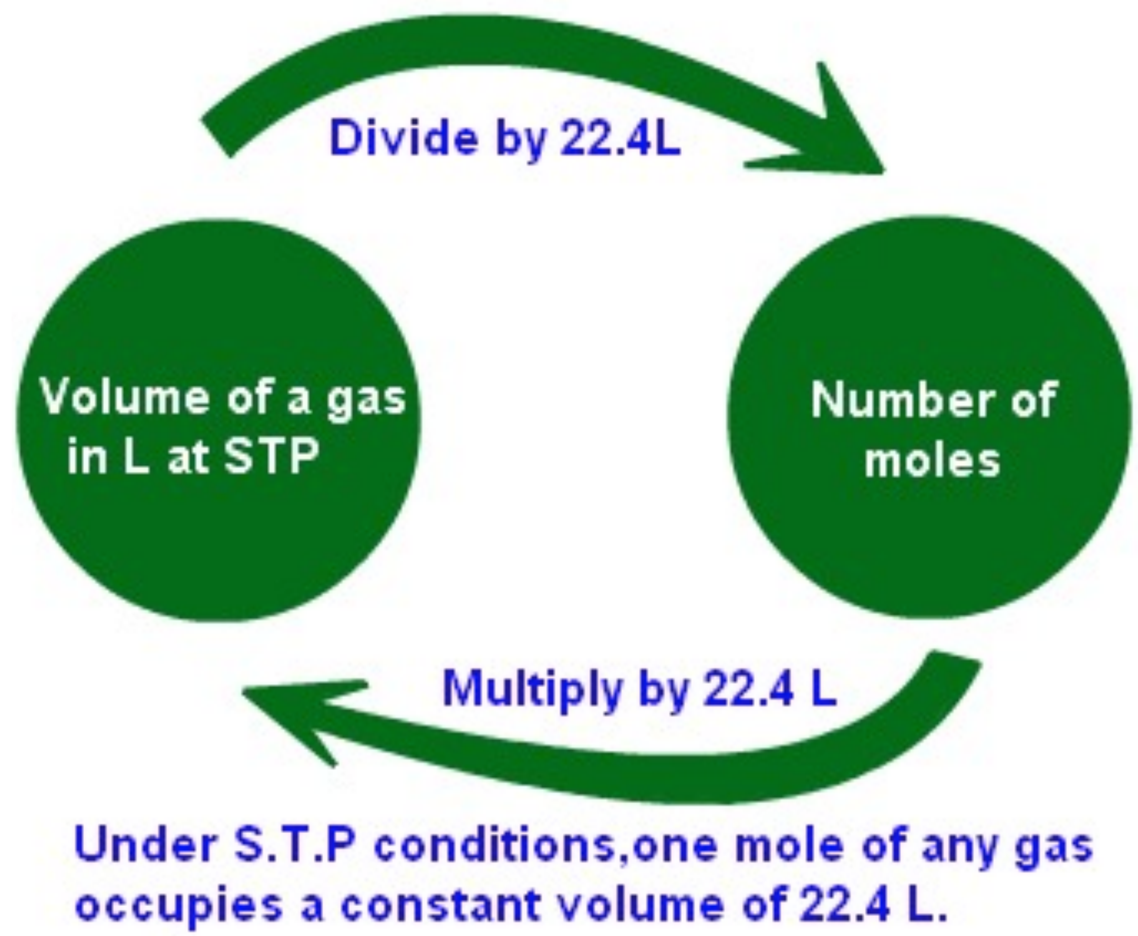
- Use the molar mass to convert between the mass and the number of moles of a substance.
- If you know the molar mass, you can convert:
  - Weight (in grams) to the moles of a substance.
  - Moles of a substance to weight.

Example

- How many moles of water are in a 355ml can of coke?
- Remember: 1ml of water is 1g.
- H_2O: 1+1+16=18 g/mole
- Use DA to solve for the number of moles of water.

The Mole-Volume Relationship

- At STP: 1 mole of a gas will occupy a volume of 22.4 liters.
- STP: Standard Temperature ($0^{\circ C}$) & Pressure (1 atm).



Your Turn

- How many moles of Oxygen are contained in 13.7 L of the gas?
- Use DA to solve.
- Remember: 1 mole = 22.4 L
- What is the volume of 37 moles of nitrogen gas?

Molar Mass from Density

- You can use the molar volume of gas to calculate the density of a gas.
- Molar Mass = density at STP x molar volume at STP.
- [grams/mole] = [grams/L] x [22.4L/mole]

10.3: Percent Composition

- % by mass: Use the molar mass and chemical formula to determine the total weight of one mole of the compound.
- Use the molar mass of the individual elements multiplied by the number of atoms of the element. Divide that number by the molar mass of the compound.

Example

- Determine the % composition of each element of H_2SO .
- Find the total molar mass of the compound: (2x1)+(1x32)+(4x16)= 98g/mol
- Determine the mass of each element & divide it by the weight of the compound:
  - Hydrogen: 2x1= (2g/mol)/(98g/mol)= 2%
  - Sulfur: 1x32= (32g/mol)/(98g/mol)=32.6%
  - Oxygen: 4x16=(64 g/mol)/(98g/mol)= 65.3%

Conversion Factors

- Use % composition to calculate the weight of any element in the mass of a compound.
- Ex: How much Hydrogen is in 11g of water (by weight)?
- 11x(2/18)=1.22g of Hydrogen

Empirical Formula

- The smallest **whole number** ratio of atoms in a compound.
- This is ratio of atoms, not the chemical formula for a particular compound.
- Example: Ethene C_2H_2 has a ratio of 1:1. Do does Styrene with C_8H_8. Both have the same ratio, but have different physical and chemical properties.

Molecular Formula

- The **exact number** of atoms in a compound.
- If there are more or less atoms in the ratio then the compound is different.
- Example: Nitric Oxide: NO. Byproduct of combustion & cardiovascular signaling molecule.
- Nitrous Oxide: N_2O: Laughing gas.

Solving for the Molecular Formula

- If you know the Empirical Formula then you can solve for the Molecular formula.
- Lets say you have a compound that tested to have a mass of 60g/mol. We know, based on the reaction, that the empirical formula is a ratio of CH_4N. What is the Molecular formula?
- Find the molar mass of the Empirical formula.
- Use that to find the ratio of the molar mass of the compound.

Computation

- The molar mass of CH_4N is 30g/mol.
- The molar mass of the compound in question is 60g/mol.
- The compound has twice the molar mass of the Empirical formula (60/30=2).
- Therefore, the Molecular formula of the compound being investigated is $C_2H_8N_2$.

Rice Lab

- Objective: Understand measurement, value, conversion, and estimation based on measures values.