# CHEMICAL EQUATIONS

- are the "recipes" in chemistry
- show the substances going into a reaction, substances coming out of the reaction, and give other information about the process

$$MgCl_2\,(aq) + 2\,AgNO_3\,(aq) \xrightarrow{\text{"yields"}} 2\,AgCl\,(s) + Mg(NO_3)_2\,(aq)$$

REACTANTS - materials that are needed fot a reaction

PRODUCTS - materials that are formed in a reaction

COEFFICIENTS - give the ratio of molecules/atoms of one substance to the others

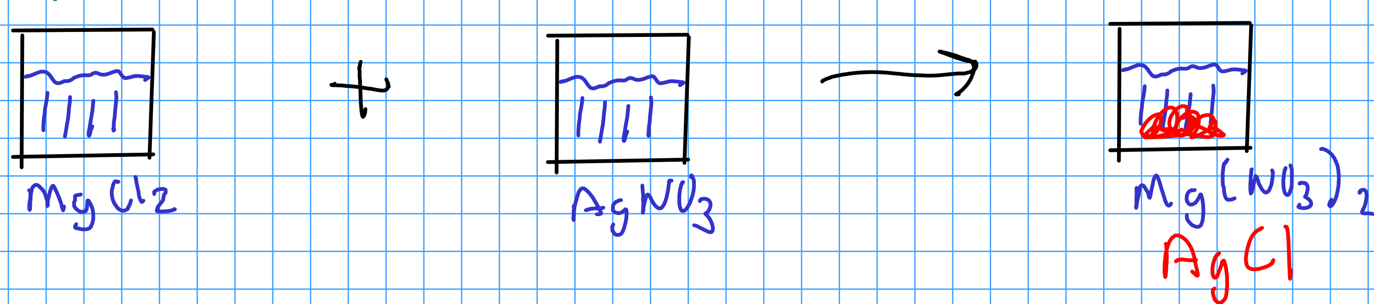
PHASE LABELS - give the physical state of a substance:

(s) -solid

(l) - liquid

(g) - gas

(aq) - aqueous. In other words, dissolved in water



# CHEMICAL EQUATIONS

$$2\,Mg(s) + O_2(g) \xrightarrow{\Delta} 2\,MgO(s)$$

REACTION CONDITIONS - give conditions necessary for chemical reaction to occur. May be:

- $\Delta$ apply hear
- catalysts - substances that will help reaction proceed faster
- other conditions, such as required temperatures

- Reaction conditions are usually written above the arrow, but may also be written below if the reaction requires several steps or several different conditions

# COEFFICIENTS

- Experimentally, we can usually determine the reactants and products of a reaction

- We can determine the proper ratios of reactants and products WITHOUT further experiments, using a process called BALANCING

- BALANCING a chemical equation is making sure the same number of atoms of each element go into a reaction as come out of it.

- A properly balanced chemical equation has the smallest whole number ratio of reactants and products.

- There are several ways to do this, but we will use a modified trial-and-error procedure.

# BALANCING

$$C_3H_8 + 5\,O_2 \rightarrow 3\,CO_2 + 4\,H_2O$$

oxygen? → 2 → 10

6 4

10

1. Pick an element. Avoid (if possible) elements that appear in more than one substance on each side of the equation.

2. Change the coefficients on substances containing this element so that the same number of atoms of the element are present on each side. CHANGE AS LITTLE AS POSSIBLE!

3. Repeat 1-2 until all elements are done.

4. Go back and quickly VERIFY that you have the same number of atoms of each element on each side, If you used any fractional coefficients, multiply each coefficient by the DENOMIMATOR of your fraction.

Use SMALLEST WHOLE NUMBER RATIOS!

# BALANCING

$$3MgCl_2 + 2Na_3PO_4 \longrightarrow Mg_3(PO_4)_2 + 6NaCl$$

$$C_2H_2 + \frac{5}{2}O_2 \longrightarrow 2CO_2 + H_2O$$

5

4 1

5

To get rid of fractional coefficients, multiply ALL coefficients by the denominator of the fraction! (x2)

$$2C_2H_2 + 5O_2 \longrightarrow 4CO_2 + 2H_2O$$

$$H_2SO_4 + 2NaOH \longrightarrow Na_2SO_4 + 2H_2O$$

Start with S, since H and O appear too many times!

After balancing Na, choose H!

# CHEMICAL CALCULATIONS - RELATING MASS AND ATOMS

$$\_\,Na_2CO_3 + \underline{2}\,HCl \rightarrow \_\,H_2O + \_\,CO_2 + \underline{2}\,NaCl$$

Chemical equations are written and balanced in terms of ATOMS and MOLECULES

- While chemical equations are written in terms of ATOMS and MOLECULES, that's NOT how we often measure substances in lab!

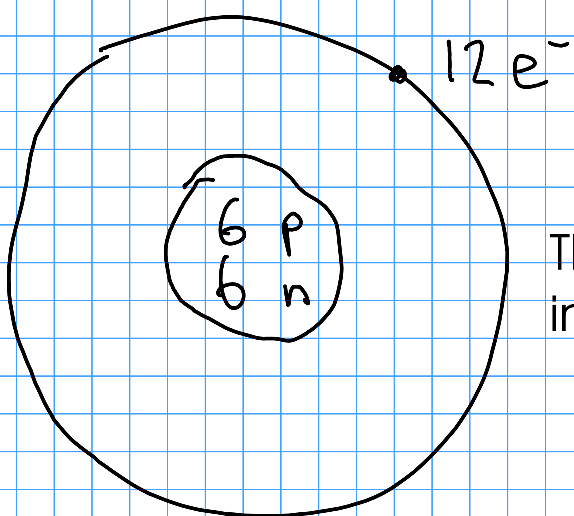
- measurements are usually MASS (and sometimes VOLUME), NOT number of atoms or molecules!

# THE MOLE CONCEPT

- A "mole" of atoms is $6.022 \times 10^{23}$ atoms

Why so big? Because atoms are so small!

- Why - in the metric dominated world of science - do we use such a strange number for quantity of atoms?




carbon-12

The mole is also defined as the number of carbon-12 atoms in exactly 12 g of carbon-12

# THE MOLE CONCEPT

- Why define the mole based on an experimentally-measured number?

- The atomic weight of an element (if you put the number in front of the unit GRAMS) is equal to the mass of ONE MOLE of atoms of that element!

Carbon (C): Atomic mass 12.01 ~~amu~~ → $\boxed{12.01\ g}$

↓

the mass of ONE MOLE of naturally-occurring carbon atoms

Magnesium (Mg): 24.31 g = the mass of ONE MOLE OF MAGNESIUM ATOMS

- So, using the MOLE, we can directly relate a mass and a certain number of atoms!

# RELATING MASS AND MOLES

- Use DIMENSIONAL ANALYSIS (a.k.a "drag and drop")

- Need CONVERSION FACTORS - where do they come from?

- We use ATOMIC WEIGHT as a conversion factor.

$Mg = 24.31$

Atomic mass

$24.31 \text{ g } Mg = 1 \underline{\text{ mol}} \; Mg$

"mol" is the abbreviation for "mole"

Example: How many moles of atoms are there in 250. g of magnesium metal?

Use the atomic weight of magnesium: $24.31 \text{ g } Mg = 1 \text{ mol } Mg$

$$250. \text{ g } \cancel{Mg} \times \frac{1 \text{ mol } Mg}{24.31 \text{ g } \cancel{Mg}} = \boxed{10.3 \text{ mol } Mg}$$

Example: You need 1.75 moles of iron. What mass of iron do you need to weigh out on thebalance? Fe: 55.85

$55.85\ g\ Fe = 1\ mol\ Fe$

$$1.75\ mol\ Fe \times \frac{55.85\ g\ Fe}{1\ mol\ Fe} = \boxed{47.7\ g\ Fe}$$

# WHAT ABOUT COMPOUNDS? FORMULA WEIGHT

Example: 25.0 g of WATER contain how many MOLES of water molecules?

$H_2O$: H : 2 x 1.008 = 2.016
O : 1 x 16.00 = 16.00

18.016 — "Formula weight" of water

Formula weight = mass of one mole of either an element OR a compound

$18.016\ g\ H_2O = 1\ mol\ H_2O$

$$25.0\ \cancel{g\ H_2O} \times \frac{1\ mol\ H_2O}{18.016\ \cancel{g\ H_2O}} = \boxed{1.39\ mol\ H_2O}$$

Formula weight goes by several names:

- For atoms, it's the same thing as ATOMIC WEIGHT
- For molecules, it;s called MOLECULAR WEIGHT
- Also called "MOLAR MASS"

Example: How many grams of ammonium carbonate do we need to weigh out to get 3.65 moles of ammonium carbonate?

$NH_4^+$ $CO_3^{2-}$

$NH_4^+$

$(NH_4)_2CO_3$ <- Formula of ammonium carbonate

N: 2 x 14.01
H: 8 x 1.008
C: 1 x 12.01
O: 3 x 16.00

96.094

$$96.094\,g\,(NH_4)_2CO_3 = 1\,mol\,(NH_4)_2CO_3$$

$$3.65\,\cancel{mol\,(NH_4)_2CO_3} \times \frac{96.094\,g\,(NH_4)_2CO_3}{1\,\cancel{mol\,(NH_4)_2CO_3}} = \boxed{351\,g\,(NH_4)_2CO_3}$$

# PERCENTAGE COMPOSITION

- sometimes called "percent composition" or "percent composition by mass"
- the percentage of each element in a compound, expressed in terms of mass

Example: Find the percentage composition of ammonium nitrate.

$NH_4NO_3$ : N : 2 x 14.01 = 28.02

$NH_4^+$ $NO_3^-$

H : 4 x 1.008 = 4.032

O : 3 x 16.00 = 48.00

These are the masses of each element in 80.052 g of ammonium nitrate!

80.052 g $NH_4NO_3$ = 1 mol $NH_4NO_4$

$$\%N: \frac{28.02\,g}{80.052\,g} \times 100\% = 35.00\%\ N$$

$$\%H: \frac{4.032\,g}{80.052\,g} \times 100\% = 5.04\%\ H$$

$$\%O: \frac{48.00\,g}{80.052\,g} \times 100\% = 59.96\%\ O$$

100.00%