LECTURE NOTES FOR GENERAL CHEMISTRY ©MM 2007

CHAPTER 3 STOICHIOMETRY

STOICHIOMETRY IS THE STUDY OF MASS RELATIONSHIPS IN CHEMICAL EQUATIONS

EACH ELEMENT HAS A CHARACTERISTIC ATOMIC MASS

EACH COMPOUND HAS A MOLECULAR MASS, EQUAL TO THE SUM OF THE MASSES OF THE ATOMS IN THE MOLECULAR FORMULA

FOR BUTANE (CH_4)

$$\begin{aligned} C &= 12.011\,\mu \\ 4 \times H &= \underline{4.032\,\mu} \\ & \quad 16.043\,\mu \end{aligned}$$

4 HYDROGENS = $4 \times 1.008\,\mu$ = $4.032\,\mu$

IONIC COMPOUNDS ARE NOT SMALL MOLECULES, BUT FORM EXTENDED CRYSTALS. WE JUST USE THE FORMULA WEIGHT

MASS % COMPOSITION

SIMPLE CALCULATION OF % = $\frac{\text{PART}}{\text{WHOLE}} \times 100$

% C IN CH_4 $\quad \frac{12.011}{16.043} = 0.74867 \times 100 = 74.867\,\%$

THE MOLE

SINCE WE KNOW THE RATIO OF ANY 2 ATOMIC MASSES, WE KNOW THAT ANY SAMPLE OF 2 ELEMENTS THAT HAS THIS MASS RATIO <u>MUST</u> HAVE THE SAME NUMBER OF EACH ATOM IN IT.

TAKE CARBON AND HYDROGEN FOR EXAMPLE

IN WHOLE NUMBERS, C = 12, H = 1

SO, 1 ATOM C AND 1 ATOM H HAVE A MASS RATIO OF $\frac{12}{1}$ 12:1

10 ATOMS C AND 10 ATOMS H $\frac{120}{10} = \frac{12}{1}$

10,000 ATOMS C AND 10,000 ATOMS H $\frac{120000}{10{,}000} = \frac{12}{1}$

ANY NUMBER OF C ATOMS AND THE SAME NUMBER OF H ATOMS MUST HAVE A MASS RATIO OF $\frac{12}{1}$

TO UNDERSTAND THE MOLE, YOU MUST LOOK AT THIS IN THE OPPOSITE SENSE

IF ANY NUMBER OF C ATOMS AND THE SAME NUMBER OF H ATOMS HAS A MASS RATIO OF 12:1

THEN ANY SAMPLE OF C AND H ATOMS THAT HAS A MASS RATIO OF 12:1 MUST HAVE THE SAME NUMBER OF EACH ATOM

ANY UNIT CAN BE USED: $\frac{\text{12 POUNDS C}}{\text{1 POUND H}}$ $\frac{\text{12 TONS C}}{\text{1 TON H}}$

$$\frac{\text{12 grams C}}{\text{1 gram H}} = \frac{6.02 \times 10^{23} \text{ C ATOMS}}{6.02 \times 10^{23} \text{ H ATOMS}} = \frac{\text{1 MOL C ATOMS}}{\text{1 MOL H ATOMS}}$$

AVOGADRO'S STATEMENT

$$12.01 \text{ g C} = 6.02 \times 10^{23} \text{ ATOMS C} = 1 \text{ MOL C ATOMS}$$

THE MOLAR MASS OF A SUBSTANCE IS ITS MOLECULAR MASS, IN GRAMS

BUTANE, CH_4 IS 16.043 g = 1 MOL CH_4

HOW MANY MOL CH_4 IS 78.01 g CH_4?

$$78.01\,g\,CH_4 \times \frac{1\,MOL\,CH_4}{16.043\,g\,CH_4} = 4.863\,MOL\,CH_4$$

287.1 MOL CH_4 IS HOW MANY g CH_4?

$$287.1\,MOL\,CH_4 \times \frac{16.043\,g\,CH_4}{1\,MOL\,CH_4} = 460.6\,g\,CH_4$$

WE USE AVOGADRO'S NUMBER ONLY WHEN THE NUMBER OF MOLECULES OR ATOMS IS INVOLVED

WHAT IS THE MASS OF A SAMPLE CONTAINING 1.50×10^{26} MOLECULES OF CH_4?

$$1.50 \times 10^{26}\,MOLECULES\,CH_4 \times \frac{1\,MOL\,CH_4}{6.02 \times 10^{23}\,MOLECULES\,CH_4} \times \frac{16.043\,g\,CH_4}{1\,MOL\,CH_4} = 4.0 \times 10^{3}\,g\,CH_4$$

WHAT IS THE MASS OF A SINGLE CH_4 MOLECULE?

$$\frac{16.043\,g\,CH_4}{1\,MOL\,CH_4} \times \frac{1\,MOL\,CH_4}{6.02 \times 10^{23}\,MOLECULES\,CH_4} = 2.66 \times 10^{-23}\,g/MOLECULE$$

MASSES OF ATOMS

THE MASS OF AN AVERAGE ATOM OF AN ELEMENT IS CALLED ITS ATOMIC WEIGHT, EXPRESSED IN ATOMIC MASS UNITS μ

SO $^{12}_{6}C$ HAS AN <u>AVERAGE</u> MASS OF 12.01

WE NEED AN AVERAGE WEIGHT BECAUSE 1.1% OF NATURALLY OCCURRING CARBON ATOMS ARE $^{13}_{6}C$

$$\text{ATOMIC WEIGHT} = \frac{\left(\begin{array}{c}\%\ \text{CARBON-12} \times \text{MASS CARBON-12} \\ + \\ \%\ \text{CARBON-13} \times \text{MASS CARBON-13}\end{array}\right)}{100}$$

$$\begin{array}{lr} 98.9\% \times 12.00 = & 11\ 87 \\ 1.1\% \times 13.00 = & 14 \\ \hline & 12\ 01 \end{array} \qquad 1201/_{100} = 12.01$$

FOR CHLORINE, Cl, THE AVERAGE ATOM WEIGHT IS CALCULATED THE SAME WAY

MASS $^{35}_{17}Cl = 34.97\mu \quad \% = 75.53$

MASS $^{37}_{17}Cl = 36.97\mu \quad \% = 24.47$

$$\text{TOMIC WEIGHT} = \frac{(75.53)(34.97) + (24.47)(36.97)}{100}$$

$$= \frac{2641.28\mu + 904.66\mu}{100} = \frac{3545.94}{100} = 35.46\mu$$

EMPIRICAL FORMULAS

- ARE THE SIMPLEST FORMULAS WE CAN WRITE USING THE SIMPLEST WHOLE-NUMBER RATIOS

MOLECULAR FORMULAS

- REPRESENT ACTUAL NUMBERS OF ATOMS IN COMPOUNDS

IF YOU CAN GET % COMPOSITION FROM A CHEMICAL FORMULA,

YOU CAN ALSO GET AN EMPIRICAL FORMULA FROM % COMPOSITION

% COMPOSITION FROM ELEMENTAL ANALYSIS

CARBON AND HYDROGEN ARE DETERMINED BY COMBUSTION, WITH WEIGHING OF THE PRODUCTS

$$C + O_2 \rightarrow CO_2$$

$$H + O_2 \rightarrow H_2O \text{ (NOT BALANCED)}$$

A SAMPLE IS BURNED IN EXCESS O_2

THE CO_2 PRODUCED IS CAPTURED AND WEIGHED

THE H_2O PRODUCED IS ALSO CAPTURED AND WEIGHED

THE MASS % COMPOSITION CAN BE CALCULATED FROM

$$g\ CO_2 \text{ PRODUCED} \rightarrow mol\ CO_2 \rightarrow g\ C \rightarrow \%\ C \text{ IN COMPOUND}$$

EXAMPLE

COMPLETE COMBUSTION OF 0.255g OF AN ALCOHOL PRODUCES 0.561g CO_2 AND 0.306g H_2O

$$g\ C = 0.561\,g\ CO_2 \times \frac{1\ MOL\ CO_2}{44.0\,g\ CO_2} \times \frac{1\ MOL\ C}{1\ MOL\ CO_2} \times \frac{12.0\,g\ C}{1\ MOL\ C} = 0.153\,g\ C$$

$$g\ H = 0.306\,g\ H_2O \times \frac{1\ MOL\ H_2O}{18\,g\ H_2O} \times \frac{2\ MOL\ H}{1\ MOL\ H_2O} \times \frac{1.01\,g\ H}{1\ MOL\ H} = 0.0343\,g\ H$$

$$\% C = \frac{0.153\,g\ C}{0.255\,g\ \text{TOTAL}} \times 100 = 60\%\,C \qquad \% H = \frac{0.0343\,g\ H}{0.255\,g\ \text{TOTAL}} = 13.5\%\ H$$

SINCE ALCOHOLS CONTAIN OXYGEN, WE ASSUME THAT THE REMAINDER IS O $100\% - (60\% + 13.5\%) = 26.5\%\ O$

EMPIRICAL FORMULAS FROM % COMPOSITION

1) CONVERT % TO MASS (USE 100g TOTAL)

$$60\%\,C \times 100\,g = 60\,g\ C$$
$$13.5\%\ H \times 100\,g = 13.5\,g\ H$$
$$26.5\%\ O \times 100\,g = 26.5\,g\ O$$

2) CONVERT MASS TO MOLES

$$60\,g\ C \times \frac{1\ \text{MOL}\ C}{12.01\,g\ C} = 5.00\ \text{MOL}\ C$$
$$13.5\,g\ H \times \frac{1\ \text{MOL}\ H}{1.01\,g\ H} = 13.4\ \text{MOL}\ H$$
$$26.5\,g\ O \times \frac{1\ \text{MOL}\ O}{16.00\,g\ O} = 1.66\ \text{MOL}\ O$$

3) "NORMALIZE" THESE MOLAR RATIOS BY DIVIDING EACH BY THE LOWEST #

$$\frac{5.00\ \text{MOL}\ C}{1.66} = 3.01\ \text{MOL}\ C$$
$$\frac{13.4\ \text{MOL}\ H}{1.66} \quad 8.07\ \text{MOL}\ H$$
$$\frac{1.66\ \text{MOL}\ O}{1.66} = 1\ \text{MOL}\ O$$

SO, C_3H_8O IS THE EMPIRICAL FORMULA

CHEMICAL EQUATIONS

ATOMS ARE NEITHER CREATED NOR DESTROYED IN A CHEMICAL REACTION

THIS IS THE LAW OF CONSERVATION OF MATTER

IN A CHEMICAL REACTION, MATTER AND MASS ARE CONSERVED

IN A CHEMICAL REACTION, COMPOUNDS OR ELEMENTS ARE CHANGED INTO DIFFERENT COMPOUNDS OR ELEMEN

A CHEMICAL REACTION IS A CHANGE IN THE IDENTITY OF COMPOUNDS, SO, THE NUMBER AND TYPE OF MOLECU MAY BE DIFFERENT ON EACH SIDE OF THE EQUATION

COMBUSTION

COMBUSTION IS THE FAMILIAR REACTION OF BURNING

IN COMBUSTION, COMPOUNDS OR ELEMENTS COMBINE WITH OXYGEN

$$C_{(s)} + O_{2\,(g)} \longrightarrow CO_{2\,(g)}$$

BALANCING EQUATIONS

CARBON AND OXYGEN CAN ALSO COMBINE TO FORM CARBON MONOXIDE, CO

WE USE SIMPLE, WHOLE NUMBER, "COEFFICIENTS" TC BALANCE EQUATIONS

BALANCING EQUATIONS

$$C_{(s)} + O_{2(g)} \longrightarrow CO_{(g)}$$

TO BALANCE OXYGEN

$$C_{(s)} + O_{2(g)} \longrightarrow 2\,CO_{(g)}$$

NOW BALANCE CARBON

$$2\,C_{(s)} + O_{2(g)} \longrightarrow 2\,CO_{(g)}$$

BURNING FUELS

METHANE, CH_4, IS ALSO KNOWN AS "NATURAL GAS"

IT IS THE SIMPLEST HYDROCARBON

HYDROCARBONS CONTAIN ONLY HYDROGEN AND CARBON

COMBUSTION OF METHANE

$$CH_{4(g)} + O_{2(g)} \longrightarrow CO_{2(g)} + H_2O_{(g)}$$

TRY TO START BALANCING WITH AN ELEMENT THAT APPEARS IN ONLY ONE SUBSTANCE ON EACH SIDE OF THE EQUATION

BALANCING H $\quad CH_4 + O_2 \rightarrow CO_2 + \underline{2}H_2O$

BALANCE O $\quad CH_4 + 2\,O_2 \rightarrow CO_2 + 2\,H_2O$

CALCULATIONS WITH CHEMICAL EQUATIONS AND MOLES ARE CALLED STOICHIOMETRY

THE COEFFICIENTS ARE THE NUMBERS IN FRONT OF THE FORMULA

THE COEFFICENTS TELL THE RATIOS OF EACH COMPONENT OF A REACTION WITH ONE ANOTHER

FOR EXAMPLE

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$$

$$2 \text{ MOLES } H_2 + 1 \text{ MOLE } O_2 \rightarrow 2 \text{ MOLES } H_2O$$

$$4 \text{ GRAMS } H_2 + 32g\ O_2 \rightarrow 36g\ H_2O$$

GRAMS-TO-MOLES OR MOLES-TO-GRAMS CALCULATIONS

1 MOLE = 1 MOLECULAR WEIGHT IN GRAMS

HOW MANY GRAMS IS 3.3 MOLES OF O_2 ?

$1 \text{ MOLE } O_2 = 32.0g \rightarrow$ CONVERSION FACTORS $\rightarrow \dfrac{1 \text{ MOL}}{32.0g}$ OR $\dfrac{32.0g}{1 \text{ MOL}}$

$$3.3 \text{ MOL} \times \frac{32.0g}{1 \text{ MOL}} = 105.6g$$

· HOW MANY MOLES O_2 IN 41.0g ?

$$41.0g \times \frac{1 \text{ MOL}}{32.0g} = 1.28 \text{ MOL}$$

MOLAR RATIO CALCULATIONS

THE EQUATION SAYS 2 MOLES H_2 REACT WITH 1MOL O_2

CONVERSION FACTORS $\dfrac{2 \text{ MOL } H_2}{1 \text{ MOL } O_2}$ OR $\dfrac{1 \text{ MOL } O_2}{2 \text{ MOL } H_2}$

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$$

TWO OTHER CONVERSION FACTORS COME FROM THIS EQUATION

$$\frac{1 \text{ MOL } O_2}{2 \text{ MOL } H_2O} \quad \text{OR} \quad \frac{2 \text{ MOL } H_2O}{1 \text{ MOL } O_2}$$

AND

$$\frac{2 \text{ MOL } H_2}{2 \text{ MOL } H_2O} \quad \text{OR} \quad \frac{2 \text{ MOL } H_2O}{2 \text{ MOL } H_2} \qquad \text{A 1:1 RATIO}$$

HOW MANY MOLES OF O_2 ARE NEEDED TO MAKE 7.0 MOL H_2O

$$7.0 \text{ MOL } H_2O \times \frac{1 \text{ MOL } O_2}{2 \text{ MOL } H_2O} \rightarrow 3.5 \text{ MOL } O_2$$

HOW MANY MOLES H_2O ARE PRODUCED WHEN 2.1 MOL H_2 REACT?

$$2.1 \text{ MOL } H_2 \times \frac{2 \text{ MOL } H_2O}{2 \text{ MOL } H_2} \rightarrow 2.1 \text{ MOL } H_2O$$

COMBINE MOLAR RATIOS AND MOLECULAR WEIGHTS

HOW MANY MOLES OF H_2O ARE FORMED FROM 100 g O_2?

$$100 \text{ g } O_2 \times \frac{1 \text{ MOL } O_2}{32.0 \text{ g}} \times \frac{2 \text{ MOL } H_2O}{1 \text{ MOL } O_2} = 6.25 \text{ MOL } H_2O$$

HOW MANY GRAMS OF H_2 ARE REQUIRED TO MAKE 100 g OF H_2O? ASSUME THAT SUFFICIENT O_2 IS PRESENT

$$100 \text{ g } H_2O \times \frac{1 \text{ MOL } H_2O}{18.0 \text{ g}} \times \frac{2 \text{ MOL } H_2}{2 \text{ MOL } H_2O} \times \frac{2.0 \text{ g}}{1 \text{ MOL } H_2} = 11 \text{ g } H_2$$

LIMITING REACTANT

SOMETIMES ONE REACTANT GETS USED UP FIRST

FOR EXAMPLE

IT TAKES ONE LEFT AND ONE RIGHT SHOE TO MAKE A PAIR

$$1 \text{ LEFT} + 1 \text{ RIGHT} \rightarrow \text{PAIR}$$

IF I HAVE 3 LEFTS AND 1 RIGHT, HOW MANY PAIRS?

$$3 \text{ LEFTS} + 1 \text{ RIGHT} \rightarrow 1 \text{ PAIR}$$

THE RIGHT SHOES LIMIT US TO 1 PAIR

FIND THE LIMITING REAGENT

$$2 H_2(g) + O_2(g) \rightarrow 2 H_2O(g)$$

IF 7.0 MOL H_2 REACT WITH 3.0 MOL O_2,
WHICH IS THE LIMITING REACTANT?
HOW MANY MOL H_2O ARE FORMED?

$$7.0 \text{ MOL } H_2 \times \frac{2 \text{ MOL } H_2O}{2 \text{ MOL } H_2} = 7.0 \text{ MOL } H_2O$$

$$3.0 \text{ MOL } O_2 \times \frac{2 \text{ MOL } H_2O}{1 \text{ MOL } O_2} = 6.0 \text{ MOL } H_2O$$

SO, O_2 IS THE LIMITING REACTANT
AND 1.0 MOL H_2 WILL BE LEFT OVER

A DIFFERENT EXAMPLE

$$Cu(s) + 2 AgNO_3(AQ) \rightarrow 2 Ag(s) + Cu(NO_3)_2(AQ)$$

$Cu = 63.6$ g/MOL

$AgNO_3 = 170$ g/MOL

$Ag = 108$ g/MOL

$Cu_{(s)} + 2AgNO_{3(AQ)} \rightarrow 2Ag_{(s)} + Cu(NO_3)_{2(AQ)}$

IF 17.0 g $AgNO_3$ REACT WITH 8.5 g Cu, HOW MANY grams OF Ag ARE PRODUCED

$$17.0\,g\,AgNO_3 \times \frac{1\,MOL\,AgNO_3}{170\,g\,AgNO_3} \times \frac{2\,MOL\,Ag}{2\,MOL\,AgNO_3} \times \frac{108\,g\,Ag}{1\,MOL\,Ag} = 10.8\,g\,Ag$$

$$8.50\,g\,Cu \times \frac{1\,MOL\,Cu}{63.6\,g\,Cu} \times \frac{2\,MOL\,Ag}{1\,MOL\,Cu} \times \frac{108\,g\,Ag}{1\,MOL\,Ag} = 28.9\,g\,Ag$$

SO, $AgNO_3$ IS THE LIMITING REAGENT

% YIELD

% YIELD IS $\frac{\text{ACTUAL YIELD}}{\text{THEORETICAL YIELD}}$

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O(g)$$

7.6 g H_2 REACT WITH EXCESS O_2 AND PRODUCE 62.0g H_2O

WHAT IS THE THEORETICAL YIELD?

WHAT IS THE PERCENT YIELD?

$$7.6\,g\,H_2 \times \frac{1\,MOL\,H_2}{2.00\,g\,H_2} \times \frac{2\,MOL\,H_2O}{2\,MOL\,H_2} \times \frac{18.0\,g\,H_2O}{1\,MOL\,H_2O} = 68.4\,g\,H_2O$$

$$\%\ YIELD = \frac{62.0\,g\,H_2O}{68.4\,g\,H_2O} = 90.6\,\%$$