Chapter 12

Stoichiometry

Stoichiometry

- Greek for "measuring elements"
- The calculations of quantities in chemical reactions based on a balanced equation.
- We can interpret balanced chemical equations several ways.

In terms of Particles

- Atom - Element
- Molecule
  - Molecular compound (non- metals)
  - or diatomic (O_2 etc.)
- Formula unit
  - Ionic Compounds (Metal and non-metal)

$2H_2 + O_2 \rightarrow 2H_2O$

- Two molecules of hydrogen and one molecule of oxygen form two molecules of water.
- $2\ Al_2O_3 \rightarrow 4Al + 3O_2$

$$2 \text{ formula units } Al_2O_3 \text{ form } 4 \text{ atoms Al and } 3 \text{ molecules } O_2$$

$2Na + 2H_2O \rightarrow 2NaOH + H_2$

$$2 \text{ atoms Na and } 2 \text{ molecules } H_2O \text{ form}$$
$$2 \text{ formula units NaOH and } 1 \text{ molecule } H_2$$

Look at it differently

- $2H_2 + O_2 \rightarrow 2H_2O$
- 2 dozen molecules of hydrogen and 1 dozen molecules of oxygen form 2 dozen molecules of water.
- 2 x (6.02 x 10^{23}) molecules of hydrogen and 1 x (6.02 x 10^{23}) molecules of oxygen form 2 x (6.02 x 10^{23}) molecules of water.
- 2 moles of hydrogen and 1 mole of oxygen form 2 moles of water.

In terms of **Moles**

- $2\,Al_2O_3 \rightarrow 4Al + 3O_2$
- The coefficients tell us how many moles of each kind

In terms of mass

- The law of conservation of mass applies
- We can check using moles
- $2H_2 + O_2 \rightarrow 2H_2O$

$$2 \text{ moles } H_2 \left(\frac{2.02 \text{ g } H_2}{1 \text{ moles } H_2}\right) = 4.04 \text{ g } H_2$$

$$1 \text{ moles } O_2 \left(\frac{32.00 \text{ g } O_2}{1 \text{ moles } O_2}\right) = \underline{32.00 \text{ g } O_2}$$

36.04 g reactants

In terms of mass

- $2H_2 + O_2 \rightarrow 2H_2O$

$$2 \text{ moles } H_2O \left(\frac{18.02 \text{ g } H_2O}{1 \text{ mole } H_2O}\right) = 36.04 \text{ g } H_2O$$

$$2H_2 + O_2 \rightarrow 2H_2O$$

$$36.04 \text{ g } (H_2 + O_2) = \quad 36.04 \text{ g } H_2O$$

Mole to mole conversions

- $2\,Al_2O_3 \rightarrow 4Al + 3O_2$
- every time we use 2 moles of Al_2O_3 we make 3 moles of O_2

$$\left(\frac{2 \text{ moles } Al_2O_3}{3 \text{ mole } O_2}\right) \quad \text{or} \quad \left(\frac{3 \text{ mole } O_2}{2 \text{ moles } Al_2O_3}\right)$$

Mole to Mole conversions

- How many moles of O_2 are produced when 3.34 moles of Al_2O_3 decompose?
- $2\,Al_2O_3 \rightarrow 4Al + 3O_2$

$$3.34 \text{ moles } Al_2O_3 \left(\frac{3 \text{ mole } O_2}{2 \text{ moles } Al_2O_3}\right) = 5.01 \text{ moles } O_2$$

Practice

- $2C_2H_2 + 5\,O_2 \rightarrow 4CO_2 + 2\,H_2O$
- If 3.84 moles of C_2H_2 are burned, how many moles of O_2 are needed?

Practice

- $2C_2H_2 + 5\ O_2 \rightarrow 4CO_2 + 2\ H_2O$
- How many moles of C_2H_2 are needed to produce 8.95 mole of H_2O?

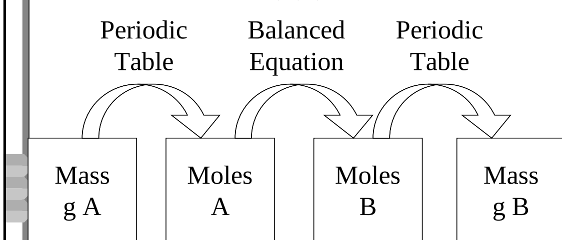
Mass in Chemical Reactions

How much do you make?

How much do you need?

We can't measure moles!!

- What can we do?
- We can convert grams to moles.
  - Periodic Table
- Then use moles to change chemicals
  - Balanced equation
- Then turn the moles back to grams.
  - Periodic table



- Decide where to start based on the units you are given
- and stop based on what unit you are asked for

Conversions

- $2C_2H_2 + 5\ O_2 \rightarrow 4CO_2 + 2\ H_2O$
- How many moles of C_2H_2 are needed to produce 8.95 g of H_2O?

Conversions

- $2C_2H_2 + 5\ O_2 \rightarrow 4CO_2 + 2\ H_2O$
- If 2.47 moles of C_2H_2 are burned, how many g of CO_2 are formed?

For example...

- If 10.1 g of Fe are added to a solution of Copper (II) Sulfate, how much solid copper would form?
- $Fe + CuSO_4 \rightarrow Fe_2(SO_4)_3 + Cu$
- $2Fe + 3CuSO_4 \rightarrow Fe_2(SO_4)_3 + 3Cu$

$$10.1 \text{ g Fe}\left(\frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}}\right)\left(\frac{3 \text{ mol Cu}}{2 \text{ mol Fe}}\right)\left(\frac{63.55 \text{ g Cu}}{1 \text{ mol Cu}}\right)$$

$$= 17.2 \text{ g Cu}$$

$2Fe + 3CuSO_4 \rightarrow Fe_2(SO_4)_3 + 3Cu$

$$0.181 \text{ mol Fe}\left(\frac{3 \text{ mol Cu}}{2 \text{ mol Fe}}\right) = 0.272 \text{ mol Cu}$$

$$0.272 \text{ mol Cu}\left(\frac{63.55 \text{ g Cu}}{1 \text{ mol Cu}}\right) = 17.3 \text{ g Cu}$$

Could have done it

$$10.1 \text{ g Fe}\left(\frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}}\right)\left(\frac{3 \text{ mol Cu}}{2 \text{ mol Fe}}\right)\left(\frac{63.55 \text{ g Cu}}{1 \text{ mol Cu}}\right)$$

$$= 17.3 \text{ g Cu}$$

- To make silicon for computer chips they use this reaction
- $SiCl_4 + 2Mg \rightarrow 2MgCl_2 + Si$
- How many moles of Mg are needed to make 9.3 g of Si?

- To make silicon for computer chips they use this reaction
- $SiCl_4 + 2Mg \rightarrow 2MgCl_2 + Si$
- 3.74 g of Mg would make how many moles of Si?

- To make silicon for computer chips they use this reaction
- $SiCl_4 + 2Mg \rightarrow 2MgCl_2 + Si$
- How many grams of $MgCl_2$ are produced along with 9.3 g of silicon?

- The U. S. Space Shuttle boosters use this reaction
- 3 Al(s) + 3 $NH_4ClO_4 \rightarrow$
  $Al_2O_3 + AlCl_3 + 3\ NO + 6H_2O$
- How much Al must be used to react with 652 g of NH_4ClO_4 ?
- How much water is produced?
- How much $AlCl_3$?

How do you get good at this?

Practice!!

Gases and Reactions

We can also change

- Liters of a gas to moles
- At STP
  - 0ºC and 1 atmosphere pressure
- At STP 22.4 L of a gas = 1 mole

For Example

- If 6.45 grams of water are decomposed, how many liters of oxygen will be produced at STP?
- $H_2O \rightarrow H_2 + O_2$
- $2H_2O \rightarrow 2H_2 + O_2$

$$6.45\text{ g } H_2O\left(\frac{1\text{ mol } H_2O}{18.02\text{ g } H_2O}\right)\left(\frac{1\text{ mol } O_2}{2\text{ mol } H_2O}\right)\left(\frac{22.4\text{ L } O_2}{1\text{ mol } O_2}\right)$$

Your Turn

- How many liters of CO_2 at STP will be produced from the complete combustion of 23.2 g C_4H_{10} ?

Example

- How many liters of CH_4 at STP are required to completely react with 17.5 L of O_2 ?
- $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

$$17.5\ \text{L}\ O_2 \left(\frac{1\ \text{mol}\ O_2}{22.4\ \text{L}\ O_2}\right)\left(\frac{1\ \text{mol}\ CH_4}{2\ \text{mol}\ O_2}\right)\left(\frac{22.4\ \text{L}\ CH_4}{1\ \text{mol}\ CH_4}\right)$$

$$= 8.75\ \text{L}\ CH_4$$

Avogadro told us

- Equal volumes of gas, at the same temperature and pressure contain the same number of particles.
- Moles are numbers of particles
- You can treat reactions as if they happen liters at a time, as long as you keep the temperature and pressure the same.

Example

- How many liters of CO_2 at STP are produced by completely burning 17.5 L of CH_4 ?
- $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

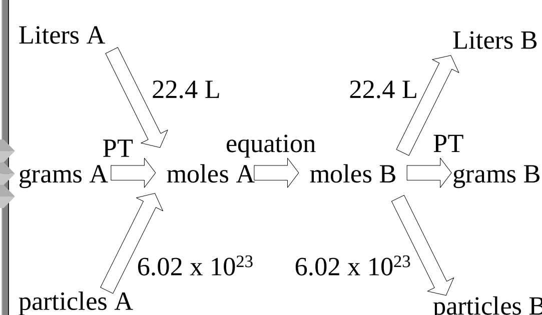
$$17.5\ \text{L}\ CH_4 \left(\frac{1\ \text{L}\ CO_2}{1\ \text{L}\ CH_4}\right) = 17.5\ \text{L}\ CO_2$$

Particles

- We can also change between particles and moles.
- 6.02×10^{23}
  - Molecules
  - Atoms
  - Formula units

Example

- If 2.8 g of C_4H_{10} are burned completely, how many water molecules will be made?



Limiting Reagent

- If you are given one dozen loaves of bread, a gallon of mustard and three pieces of salami, how many salami sandwiches can you make?
- The limiting reagent is the reactant you run out of first.
- The excess reagent is the one you have left over.
- The limiting reagent determines how much product you can make

How do you find out?

- Do two stoichiometry problems.
- The one that makes the least product is the limiting reagent.
- For example
- Copper reacts with sulfur to form copper (I) sulfide. If 10.6 g of copper reacts with 3.83 g S how much product will be formed?

- If 10.6 g of copp… …83 g S. How many gram… formed?
- 2Cu + S $\rightarrow$ C…

Cu is Limiting Reagent

$10.6\text{ g Cu}\left(\frac{1\text{ mol}}{63.55\text{g C}\ldots}\right) \ldots \left(\frac{\ldots Cu_2S}{\ldots ol\ Cu_2S}\right)$

$13.3\text{ g } Cu_2S$

$$3.83\text{ g S}\left(\frac{1\text{ mol S}}{32.06\text{g S}}\right)\left(\frac{1\text{ mol } Cu_2S}{1\text{ mol S}}\right)\left(\frac{159.16\text{ g } Cu_2S}{1\text{ mol } Cu_2S}\right)$$

$= 19.0\text{ g } Cu_2S$

How much excess reagent?

- Use the limiting reagent to find out how much excess reagent you used
- Subtract that from the amount of excess you started with

- $Mg(s) + 2\ HCl(g) \rightarrow MgCl_2(s) + H_2(g)$
- If 4.87 mol of magnesium and 9.84 mol of HCl gas are reacted, how many moles of gas will be produced?
- What is the limiting reagent?

- $Mg(s) + 2\ HCl(g) \rightarrow MgCl_2(s) + H_2(g)$
- If 4.87 mol of magnesium and 9.84 mol of HCl gas are reacted, how many moles of gas will be produced?
- How much excess reagent remains?

- If 10.3 g of aluminum are reacted with 51.7 g of $CuSO_4$ how much copper will be produced?
- How much excess reagent will remain?

Yield

- The amount of product made in a chemical reaction.
- There are three types
- Actual yield- what you get in the lab when the chemicals are mixed
- Theoretical yield- what the balanced equation tells you you should make.
- Percent yield = $\dfrac{\text{Actual}}{\text{Theoretical}} \times 100\ \%$

Example

- 6.78 g of copper is produced when 3.92 g of Al are reacted with excess copper (II) sulfate.
- $2Al + 3\ CuSO_4 \rightarrow Al_2(SO_4)_3 + 3Cu$
- What is the actual yield?
- What is the theoretical yield?
- What is the percent yield?
- If you had started with 9.73 g of Al, how much copper would you expect?

Details

- Percent yield tells us how "efficient" a reaction is.
- Percent yield can not be bigger than 100 %.

Energy in Chemical Reactions

How Much?

In or Out?

Energy

- Energy is measured in Joules or calories
- Every reaction has an energy change associated with it
- Exothermic reactions release energy, usually in the form of heat.
- Endothermic reactions absorb energy
- Energy is stored in bonds between atoms
- Making bonds gives energy
- Breaking bonds takes energy

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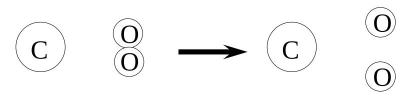
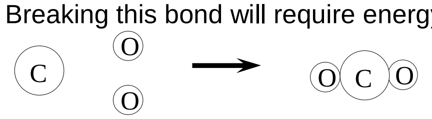
In terms of bonds



Breaking this bond will require energy



Making these bonds gives you energy

In this case making the bonds gives you more energy than breaking them

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Exothermic

- The products are lower in energy than the reactants
- Releases energy
- Often release heat

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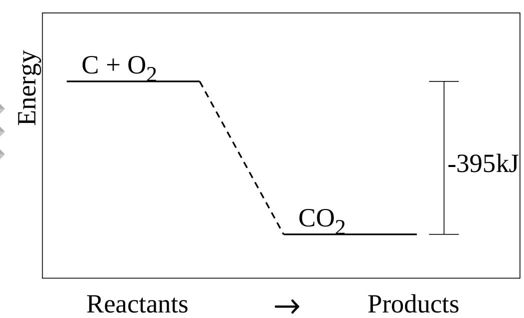
$C + O_2 \rightarrow CO_2 + 395\ kJ$



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Endothermic

- The products are higher in energy than the reactants
- Absorbs energy
- Absorb heat

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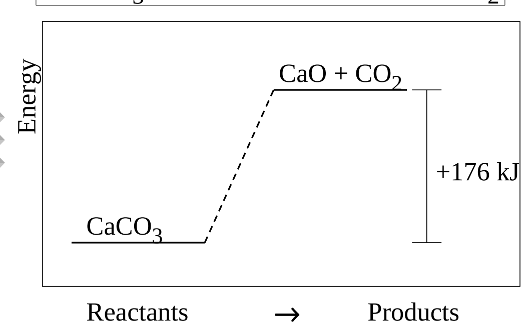
$CaCO_3 + 176\ kJ \rightarrow CaO + CO_2$



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Chemistry Happens in

- MOLES
- An equation that includes energy is called a thermochemical equation
- $CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O + 802.2\ kJ$
- 1 mole of CH_4 makes 802.2 kJ of energy.
- When you make 802.2 kJ you make 2 moles of water

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$CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O + 802.2\ kJ$

- If 10. 3 grams of CH_4 are burned completely, how much heat will be produced?

$$\mathbf{10.\ 3\ g\ CH_4 \left(\frac{1\ mol\ CH_4}{16.05\ g\ CH_4}\right)\left(\frac{802.2\ kJ}{1\ mol\ CH_4}\right)}$$

=515 kJ

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$CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O + 802.2\ kJ$

- How many liters of O_2 at STP would be required to produce 23 kJ of heat?
- How many grams of water would be produced with 506 kJ of heat?

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Calorimetry

- Measuring heat.
- Use a calorimeter.
- Two kinds
- Constant pressure calorimeter (called a coffee cup calorimeter)
- An insulated cup, full of water.
- The specific heat of water is 1 cal/g$^{\circ}$C
- heat = specific heat x m x ΔT

Example

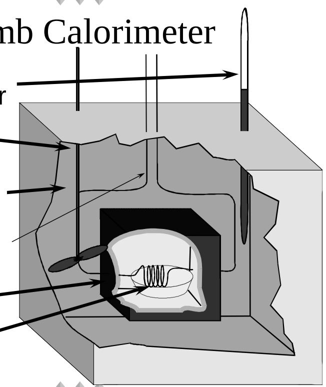
- A chemical reaction is carried out in a coffee cup calorimeter. There are 75.8 g of water in the cup, and the temperature rises from 16.8 $^{\circ}$C to 34.3 $^{\circ}$C. How much heat was released?

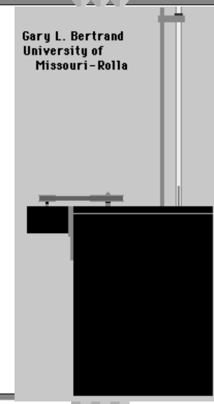
Calorimetry

- Second type is called a bomb calorimeter.
- Material is put in a container with pure oxygen. Wires are used to start the combustion. The container is put into a container of water.
- The heat capacity of the calorimeter is known and/or tested. (cal/ $^{\circ}$C)
- Multiply temperature change by the heat capacity to find heat

Bomb Calorimeter

- thermometer
- stirrer
- full of water
- ignition wire
- Steel bomb
- sample





Heats of Reaction

Enthalpy

- The heat content a substance has at a given temperature and pressure
- Can't be measured directly because there is no set starting point
- The reactants start with a heat content
- The products end up with a heat content
- So we can measure how much heat content changes
- Called change in enthalpy

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Enthalpy

- Symbol is H
- Change in enthalpy is ΔH
- delta H
- If heat is released the heat content of the products is lower
- ΔH is negative (exothermic)
- If heat is absorbed the heat content of the products is higher
- ΔH is positive (endothermic)

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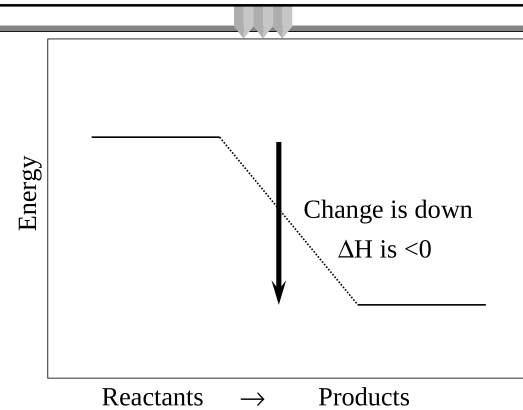

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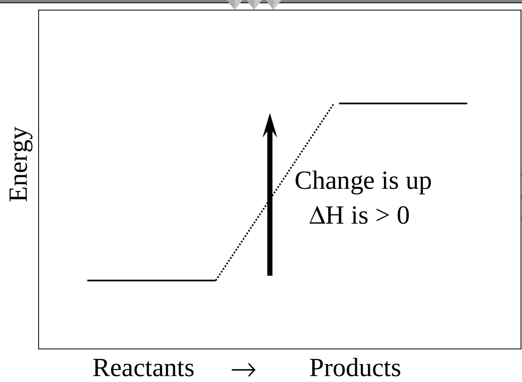

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Heat of Reaction

- The heat that is released or absorbed in a chemical reaction
- Equivalent to ΔH
- $C + O_2(g) \rightarrow CO_2(g) + 393.5$ kJ
- $C + O_2(g) \rightarrow CO_2(g)$ $\Delta H = -393.5$ kJ
- In thermochemical equation it is important to say what state
- $H_2(g) + \frac{1}{2} O_2(g) \rightarrow H_2O(g)$ $\Delta H = -241.8$ kJ
- $H_2(g) + \frac{1}{2} O_2(g) \rightarrow H_2O(l)$ $\Delta H = -285.8$ kJ

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Heat of Combustion

- The heat from the reaction that completely burns 1 mole of a substance
- $C_2H_4 + 3\,O_2 \rightarrow 2\,CO_2 + 2\,H_2O$
- $C_2H_6 + O_2 \rightarrow CO_2 + H_2O$
- $2\,C_2H_6 + 7\,O_2 \rightarrow 2\,CO_2 + 6\,H_2O$
- $C_2H_6 + (7/2)\,O_2 \rightarrow CO_2 + 3\,H_2O$

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Standard Heat of Formation

- The ΔH for a reaction that produces 1 mol of a compound from its elements at standard conditions
- Standard conditions 25°C and 1 atm.
- Symbol is ΔH^o_f
- The standard heat of formation of an element is 0
- This includes the diatomics

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What good are they?

- There are tables (pg. 190) of heats of formations
- For most compounds it is negative
  - Because you are making bonds
  - Making bonds is exothermic
- The heat of a reaction can be calculated by subtracting the heats of formation of the reactants from the products

$\Delta H = \Delta H^o_f(\text{products}) - \Delta H^o_f(\text{reactants})$

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Examples

- $CH_4(g) + 2\,O_2(g) \rightarrow CO_2(g) + 2\,H_2O(g)$

ΔH^o_f **CH_4 (g) = -74.86 kJ**

ΔH^o_f **O_2(g) = 0 kJ**

ΔH^o_f **CO_2(g) = -393.5 kJ**

ΔH^o_f **H_2O(g) = -241.8 kJ**

- **ΔH= [-393.5 + 2(-241.8)]-[-74.86 +2 (0)]**
- **ΔH= -802.2 kJ**

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Examples

- $2\ SO_3(g) \rightarrow 2SO_2(g) + O_2(g)$

Why Does It Work?

- If $H_2(g) + 1/2\ O_2(g) \rightarrow H_2O(l)$ $\Delta H = -285.5$ kJ
- then
  $H_2O(g) \rightarrow H_2(g) + 1/2\ O_2(l)$ $\Delta H = +285.5$ kJ
- If you turn an equation around, you change the sign
- $2\ H_2O(g) \rightarrow 2\ H_2(g) + O_2(l)$ $\Delta H = +571.0$ kJ
- If you multiply the equation by a number, you multiply the heat by that number.

Why does it work?

- You make the products, so you need their heats of formation
- You “unmake” the reactants so you have to subtract their heats.