SINGLE REPLACEMENT REACTIONS

$$A + BC \longrightarrow AC + B$$

One element, usually a metal, replaces another element in a compound. This forms a new compound and leaves behind a new free element!

example: $\underset{0}{Cu}(s) + 2\underset{+1}{Ag}NO_3(aq) \longrightarrow \underset{+2}{Cu}(NO_3)_2(aq) + 2\underset{0}{Ag}(s)$

Copper loses electrons, goes from 0 charge to +2 charge!

Silver gains electrons, goes from +1 charge to 0 charge!

... but just because you combine an element and a compound doesn't mean that a reaction will occur. Some combinations react, some don't!

- Whether a reaction occurs depends on how easily the replacing and replaced elements lose electrons. An atom that loses electrons more easily will end up in IONIC form (in other words, in the compound). An atom that loses electrons less easily will end up as a free element.

- We say that an atom that loses elecrons more easily that another is MORE ACTIVE than the other element. But how would you get information about ACTIVITY?

A single replacement reaction is an example of a reaction where ELECTRON TRANSFER is a driving force. Electron transfer reactions are generally called OXIDATION-REDUCTION reactions.

ACTIVITY SERIES

- comes from experiental data. It's a list of elements in order of their ACTIVITY - more active elements are higher in the series!

A sample activity series

Activity

Sodium Na^+

Magnesium Mg^{2+}

Aluminum Al^{3+}

Zinc Zn^{2+}

Iron Fe^{2+}

Lead Pb^{2+}

Hydrogen H^+

Copper Cu^{2+}

Silver Ag^+

Gold Au^{3+}

Very active metals will replace hydrogen in acids AND in water!

Metals more active than hydrogen will replace hydrogen in acids!

These metals are unreactive to most acids!

PREDICTING SINGLE REPLACEMENT REACTIONS

$$Pb(s) + 2HCl(aq) \rightarrow PbCl_2(s) + H_2(g)$$

Lead is MORE ACTIVE than hydrogen, so we would expect lead to replace hydrogen in this reaction.

$$Pb(NO_3)_2(aq) + Zn(s) \rightarrow Zn(NO_3)_2(aq) + Pb(s)$$

Zinc is MORE ACTIVE than lead, so we would expect zinc to replace lead in this reaction.

$Ag(s) + H_2SO_4(aq) \rightarrow$ NO REACTION

Silver is LESS ACTIVE than hydrogen, so we would NOT expect it to be able to replace hydrogen in a single replacement reaction.

$$Mg(s) + ZnSO_4(aq) \rightarrow MgSO_4(aq) + Zn(s)$$

Since magnesium is MORE ACTIVE than zinc, we would expect magnesium to replace zinc in this reaction.

Activity

Sodium	Na^+
Magnesium	Mg^{2+}
Aluminum	Al^{3+}
Zinc	Zn^{2+}
Iron	Fe^{2+}
Lead	Pb^{2+}
Hydrogen	H^+
Copper	Cu^{2+}
Silver	Ag^+
Gold	Au^{3+}

CHEMICAL CALCULATIONS - RELATING MASS AND ATOMS

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

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!

Na_2CO_3 solid

HCl solution

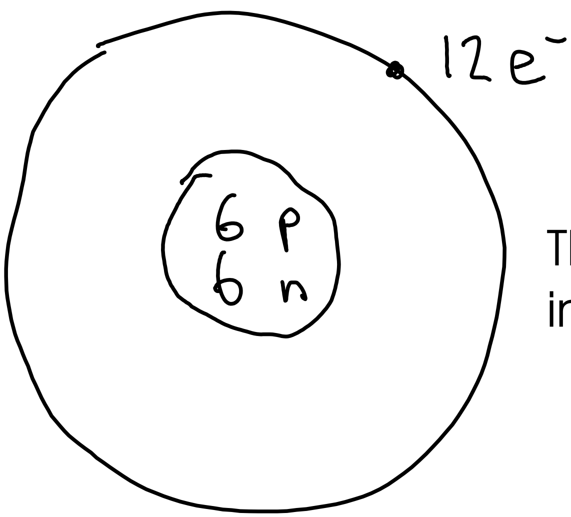
... so how do we relate atoms and molecules with things we routinely measure in lab - like grams and milliliters?

THE MOLE CONCEPT

- A "mole" of atoms is 6.022×10^{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.

$$Mg: 24.31 \quad \Big| \quad 24.31\ g\ Mg = 1\ \underline{mol}\ Mg$$

"mol" is the abbreviation for "mole"

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

$$24.31\ g\ Mg = mol\ Mg$$

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

Example: You need 1.75 moles of iron. What mass of iron do you need to weigh out on the balance?

Use the atomic weight of iron (55.85 g/mol) as a conversion factor relating mass and moles.

$$55.85\,g\,Fe = mol\,Fe$$

$$1.75\,\cancel{mol\,Fe} \times \frac{55.85\,g\,Fe}{\cancel{mol\,Fe}} = \boxed{97.7\,g\,Fe}$$

WHAT ABOUT COMPOUNDS? FORMULA WEIGHT

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

(H_2O)

H_2O H : $2 \times 1.008 = 2.016$

O : $1 \times 16.00 = 16.00$

18.016 — FORMULA WEIGHT of water

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

$18.016\ g\ H_2O = mol\ H_2O$

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

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^+

$(NH_4)_2CO_3$

N: 2 x 14.01

H: 8 x 1.008

C: 1 x 12.01

O: 3 x 16.00

96.094 <-- Formula weight of ammonium carbonate

Use the formula weight as a conversion factor

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

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

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_4^+ NO_3^-

$\underline{NH_4NO_3}$:

$N: 2 \times 14.01 = 28.02$

$H: 4 \times 1.008 = 4.032$

$O: 3 \times 16.00 = 48.00$

$80.052\ g\ NH_4NO_3 = 1\ mol\ NH_4NO_3$

$\%N: \frac{28.02\,g\,N}{80.052\,g\,tot} \times 100\% = 35.0\%\,N$

$\%H: \frac{4.032\,g\,H}{80.052\,g\,tot} \times 100\% = 5.0\%\,H$

$\%O: \frac{48.00\,g\,O}{80.052\,g\,tot} \times 100\% = 60.0\%\,O$

These percentages should sum to 100%, but you may have a little roundoff error.

sum to 100%

So far, we have

ch8

- looked at how to determine the composition by mass of a compound from a formula
- converted from MASS to MOLES (related to the number of atoms/molecules)
- converted from MOLES to MASS

Are we missing anything?

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- What about SOLUTIONS, where the desired chemical is not PURE, but found DISSOLVED IN WATER?
- How do we deal with finding the moles of a desired chemical when it's in solution?

MOLAR CONCENTRATION

- unit: MOLARITY (M): moles of dissolved substance per LITER of solution

$$M = \text{MOLARITY} = \frac{\text{moles of solute}}{\text{L solution}}$$

$$6.0\ M\ HCl\ \text{solution:}\ \frac{6.0\ \text{mol}\ HCl}{L}$$

There are 6.0 moles of hydrochloric acid in each liter of this solution, so you can write this relationship another way:

$$6.0\ \text{mol}\ HCl = 1\ L$$

If you have 0.250 L (250 mL) of 6.0 M HCl, how many moles of HCl do you have?

$$6.0\ \text{mol}\ HCl = L$$

$$0.250\ L \times \frac{6.0\ \text{mol}\ HCl}{L} = \boxed{1.5\ \text{mol}\ HCl}$$

If you need 0.657 moles of hydrochloric acid, how many milliliters of 0.0555 M HCl do you need to measure out?

$0.0555 \text{ mol HCl} = \text{L} \qquad \text{mL} = 10^{-3}\text{L}$

$0.657 \text{ mol} \cancel{\text{HCl}} \times \dfrac{\cancel{\text{L}}}{0.0555 \text{ mol} \cancel{\text{HCl}}} \times \dfrac{\text{mL}}{10^{-3}\cancel{\text{L}}} = \boxed{11800 \text{ mL}}$

This is a volume that's not practical in a lab setting. Maybe we should use a more concentrated solution?

What if we used 6.00 M HCl?

$6.00 \text{ mol HCl} = \text{L} \qquad \text{mL} = 10^{-3}\text{L}$

$0.657 \text{ mol} \cancel{\text{HCl}} \times \dfrac{\cancel{\text{L}}}{6.00 \text{ mol} \cancel{\text{HCl}}} \times \dfrac{\text{mL}}{10^{-3}\cancel{\text{L}}} = \boxed{110. \text{ mL}}$

This is a more reasonable volume to measure in a laboratory.

If you're preparing a solution by dissolving a solid in water, you can easily calculate the molarity of the solution. How?

Just find the number of moles of solid you dissolved, then divide by the volume of the solution (expressed in liters!)

What is the molarity of a solution made by dissolving 3.50 g of NaCl in enough water to make 250. mL of solution?

$$M = \frac{\text{mol NaCl}}{\text{L solution}}$$

Find moles NaCl by using the formula weight

Na: 1 x 22.99

Cl: 1 x 35.45

58.44 g NaCl = mol NaCl

$$3.50\text{ g NaCl} \times \frac{\text{mol NaCl}}{58.44\text{ g NaCl}} = 0.059890\text{ mol NaCl}$$

$$M = \frac{0.059890\text{ mol NaCl}}{0.250\text{ L}}$$

$$= \boxed{0.240\ \underline{M}\text{ NaCl}}$$

Find volume of solution

$$\text{mL} = 10^{-3}\text{ L}$$

$$250\text{ mL} \times \frac{10^{-3}\text{ L}}{\text{mL}} = 0.250\text{ L}$$

A few more examples... Use FORMULA WEIGHT when relating mass and moles

You have a 250.g bottle of silver(I) chloride (AgCl). How many moles of AgCl do you have?

AgCl : Ag : 1x 107.9
Cl : 1x 35.45

143.35 g AgCl = mol AgCl

$$250.\,g\ AgCl \times \frac{mol\ AgCl}{143.35\,g\ AgCl} = \boxed{1.74\ mol\ AgCl}$$

How many moles of NaOH are present in 155 mL of 1.50 M NaOH?

When relating moles and VOLUME, we need to use CONCENTRATION (usually MOLARITY - M)

1.50 mol NaOH = L   mL = 10^{-3} L

$$155\,\cancel{mL} \times \frac{10^{-3}\,\cancel{L}}{\cancel{mL}} \times \frac{1.50\ mol\ NaOH}{\cancel{L}} = \boxed{0.233\ mol\ NaOH}$$

End of material for test 3
Summer 2009,
Test date 7/21/09