> You always pass failure on the way to success.
> --Mickey Rooney

Making sense out of Chemical Reactions

Wait! Don't throw this away in disgust! It really is possible to do it! Read, think and learn:

We have arbitrarily (and conveniently) divided inorganic chemical reactions into THREE large categories. The bottom line is that you need to learn how to recognize, write and/or complete, and balance these reactions. They are the "stuff" of which chemistry is made.

Precipitation reactions

How to recognize: the ones you will see in this course involve TWO compounds as possible reactants.

e.g. sodium sulfate + barium chloride → ??????

$Na_2SO_4 + BaCl_2 \rightarrow$??????

is possibly a precipitation reaction.

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iron + hydrochloric acid → ????????

$Fe + HCl \rightarrow$?????????

is definitely not.

How to complete: try switching the ions and check products against the solubility rules; if at least one product is insoluble, you have a reaction; if both possible products are soluble, you have a dud (i.e., no reaction)

e.g. $Na_2SO_4 + BaCl_2 \rightarrow NaCl + BaSO_4$ (not balanced!!!)

according to the solubility rules, $BaSO_4$ is insoluble. So there is a reaction.

$NaNO_3 + BaCl_2 \rightarrow Ba(NO_3)_2 + NaCl$ (not balanced!!!)

according to the solubility rules, both possible products are soluble. NO reaction will occur. Such a problem is best written (in molecular form) as:

$NaNO_3 + BaCl_2 \rightarrow$ N.R.

How to balance: make sure you have the correct formulas first by checking the charges of the ions. Then AND ONLY THEN use coefficients in front of the formulas to get the same number of atoms on both sides. **DO NOT ALTER FORMULAS TO FORCE A REACTION TO BALANCE**.

e.g. $Na_2SO_4 + BaCl_2 \rightarrow 2\ NaCl + BaSO_4$

How to write the net-ionic version: re-write the reaction, breaking up all soluble substances into ions, cancel spectators and then check balance.

e.g. $2\ Na^+ + SO_4^{2-} + Ba^{2+} + 2\ Cl^- \rightarrow 2\ Na^+ + 2\ Cl^- + BaSO_4$

--OR--

$SO_4^{2-} + Ba^{2+} \rightarrow BaSO_4$

Acid Reactions

How to recognize: at this point, reactants must include one of the five common acids you are supposed to know and a hydroxide compound (a base) or a metal oxide or a metal carbonate These are all compounds, so be careful not to confuse this with a precipitation reaction.

e.g. hydrochloric acid + sodium hydroxide → ?????

$HCl + NaOH \rightarrow$?????

is an acid-base reaction (neutralization)

$HNO_3 + NaCl \rightarrow$??????

is not, since there is no base present.

How to complete: all common acid-base reactions you will be seeing at this point produce water and an ionic compound or salt. The salt may be soluble or insoluble.

e.g. $HCl + NaOH \rightarrow H_2O + NaCl$

note that this result is obtained the same way as the products for a precipitation reaction--switch the ions. This example is already balanced.

How to balance: do this in the same way as you do precipitation reactions

How to write the net-ionic version: follow the same procedure as for precipitation reactions, bearing in mind that all acids and bases are electrolytes and water is <u>not</u>. Electrolytes break up into ions in solution.

e.g. $H^+ + Cl^- + Na^+ + OH^- \rightarrow H_2O + Na^+ + Cl^-$

--OR--

$H^+ + OH^- \rightarrow H_2O$

Other sequences:

Remember, acids react with metal oxides to give water and a salt:

e.g. $HCl + MgO \rightarrow MgCl_2 + H_2O$ (not balanced)

$2\ HCl + MgO \rightarrow MgCl_2 + H_2O$ (balanced molecular)

Since metal oxides are solids, the net-ionic equation must show this:

$2\ H^+ + 2\ Cl^- + MgO \rightarrow Mg^{2+} + 2\ Cl^- + H_2O$ (total ionic--not balanced)

--OR--

$2\ H^+ + MgO \rightarrow Mg^{2+} + H_2O$ (balanced net-ionic)

For acids reacting with metal carbonates, the script is similar, with the addition of carbon dioxide as a third product. Details vary depending on whether the carbonate is soluble (like Li_2CO_3) or not.

e.g. $H_2SO_4 + Li_2CO_3 \rightarrow Li_2SO_4 + H_2O + CO_2$ (balanced molecular)

$2\ H^+ + SO_4^{2-} + 2\ Li^+ + CO_3^{2-} \rightarrow 2\ Li^+ + SO_4^{2-} + H_2O + CO_2$

$2\ H^+ + CO_3^{2-} \rightarrow H_2O + CO_2$ (balanced net-ionic)

Redox Reactions

How to recognize: if it doesn't look like precipitation or acid, it is possibly redox of some kind. There are several possibilities:

1) synthesis from elements

$Na + Cl_2 \rightarrow$???????

2) decomposition of a simple compound

$HgO \rightarrow$??????

3) displacement

a. metals with acids

$Zn + HCl \rightarrow$?????

b. metals with metal compounds

$Zn + Cu(NO_3)_2 \rightarrow$??????

c. halogens with halide compounds

$Cl_2 + KI \rightarrow$????

Note that all of these follow a simple pattern: they involve *an element and a compound*

4) OTHER (ugh!) (this includes anything else)

How to complete:

1 & 2) Only very obvious synthesis and decomposition reactions will be given or else very generous hints will also be given.

e.g. $Na + Cl_2 \rightarrow NaCl$ (not balanced)

$HgO \rightarrow Hg + O_2$ (not balanced)

3) You must check the metal activity series or the halogen family before deciding if a reaction will occur. Remember, in each case, the element which is a possible reactant must be higher in the list than the part of the compound it is supposed to displace. How do you know what will be displaced? "Like displaces like" (i.e., metals displace metals, positive displaces positive, etc.)

e.g. $Zn + HCl \rightarrow ZnCl_2 + H_2$ (not balanced)

This works since zinc is higher than hydrogen in the activity series. Note that in compounds they are both positive.

$Cu + HCl \rightarrow$ N.R.

This will not work since copper is below hydrogen in the activity series.

$Cl_2 + KI \rightarrow KCl + I_2$ (not balanced)

This works since chlorine is above iodine in the halogen family.

$I_2 + KCl \rightarrow$ N.R.

This will not work since iodine is below chlorine in the halogen column.

4) OTHER are so complicated that we will have to give you the products so you needn't worry about figuring them out by yourself.

How to balance: this is done in the usual way for all redox reactions except OTHER. See the end of this handout for examples.

How to write the net-ionic version: follow the usual procedure, keeping in mind that elements (like halogens) are not electrolytes and should not be broken up. *Be sure to check that the charge is balanced as well as the atoms.*

e.g. $Zn + 2\,H^+ + 2\,\cancel{Cl^-} \rightarrow H_2 + Zn^{2+} + 2\,\cancel{Cl^-}$

$Cl_2 + \cancel{K^+} + 2\,I^- \rightarrow I_2 + 2\,Cl^- + \cancel{K^+}$

Balancing non-trivial redox reactions

Most of these reactions occur between ions in solution. The solution is usually either acidic or basic. The good news is that if you follow the method outlined (same one as in your text) you will end up with a balanced net-ionic reaction. You can decide for yourself what the bad news is.

example: permanganate ions react with Cl^- in acidic solution to produce Mn^{2+} and Cl_2

1. write "skeleton" reaction:

$$MnO_4^- + Cl^- \rightarrow Mn^{2+} + Cl_2 \quad (H^+)$$

fine print: the H^+ tells you the solution is acidic

2. break up the reaction into halves, each containing an element that is changing, and balance those elements if needed:

$$MnO_4^- \rightarrow Mn^{2+}$$

$$2\,Cl^- \rightarrow Cl_2$$

3. Since these reactions occur in solution, use H_2O to balance oxygen atoms if needed:

$$MnO_4^- \rightarrow Mn^{2+} + 4\,H_2O$$

$$2\,Cl^- \rightarrow Cl_2$$

4. Use H^+ to balance hydrogens if needed:

$$8\,H^+ + MnO_4^- \rightarrow Mn^{2+} + 4\,H_2O$$

$$2\,Cl^- \rightarrow Cl_2$$

5. Use electrons to balance charges:

$$5\,e^- + 8\,H^+ + MnO_4^- \rightarrow Mn^{2+} + 4\,H_2O$$

$$2\,Cl^- \rightarrow Cl_2 + 2\,e^-$$

6. Multiply half-reactions to equalize electrons (find LCM):

$$2\,(5\,e^- + 8\,H^+ + MnO_4^- \rightarrow Mn^{2+} + 4\,H_2O)$$

$$5\,(2\,Cl^- \rightarrow Cl_2 + 2\,e^-)$$

7. Add half-reactions together, simplifying:

$$16\ H^+ + 2\ MnO_4^- + 10\ Cl^- \rightarrow 2\ Mn^{2+} + 8\ H_2O + 5\ Cl_2$$

Whew.

Question: What do I do if the reaction takes place in basic solution?

Answer: exactly the same thing and then add to each side as many OH^- as there are H^+. Combine H^+ and OH^- on the same side to make H_2O and simplify. Altering the previous problem (just for make-believe):

$$16\ OH^- + 16\ H^+ + 2\ MnO_4^- + 10\ Cl^- \rightarrow 2\ Mn^{2+} + 8\ H_2O + 5\ Cl_2 + 16\ OH^-$$

$$16\ H_2O + 2\ MnO_4^- + 10\ Cl^- \rightarrow 2\ Mn^{2+} + 8\ H_2O + 5\ Cl_2 + 16\ OH^-$$

$$8\ H_2O + 2\ MnO_4^- + 10\ Cl^- \rightarrow 2\ Mn^{2+} + 5\ Cl_2 + 16\ OH^-$$