5.2 Balancing chemical equations

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

s - solid

l - liquid

g - gas

aq - aqueous

Another example: do oxygen last…

$$C_2H_6 \quad + \quad O_2 \quad \rightarrow \quad CO_2 \quad + \quad H_2O$$

Case without a "1" stoichiometric coef

A stoichiometric coefficient is the number in front of each chemical species.

$C_8H_{18}\,(l) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

$C_8H_{18}\,(l) + O_2(g) \rightarrow 8CO_2(g) + H_2O(g)$

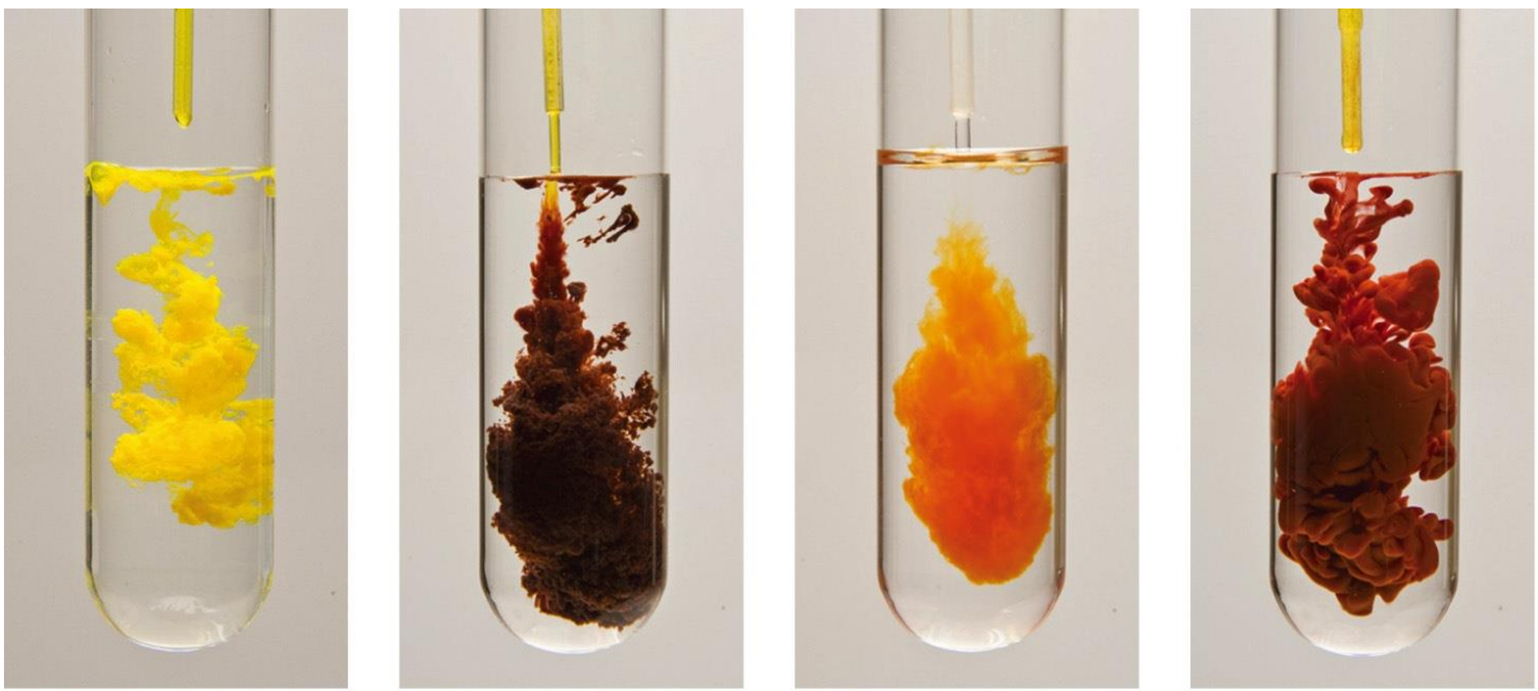
$C_8H_{18}\,(l) + O_2(g) \rightarrow 8CO_2(g) + 9H_2O(g)$

$C_8H_{18}\,(l) + 25/2\,O_2(g) \rightarrow 8CO_2(g) + 9H_2O(g)$

$2C_8H_{18}\,(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$

5.4 Precipitation Reactions

- A reaction where an insoluble solid (precipitate) forms and falls out of solution



$PbCrO_4$ from $Pb(NO_3)_2$ and K_2CrO_4

PbS from $Pb(NO_3)_2$ and $(NH_4)_2S$

$Fe(OH)_3$ from $FeCl_3$ and NaOH

Ag_2CrO_4 from $AgNO_3$ and K_2CrO_4

Soluble	Exceptions
Ammonium compounds (NH^+_4)	None
Lithium compounds (Li^+)	None
Sodium compounds (Na^+)	None
Potassium compounds (K^+)	None
Nitrates (NO^-_3)	None
Perchlorates (ClO^-_4)	None
Acetates ($CH_3CO^-_2$)	None
Chlorides (Cl^-)	Ag^+, Hg_2^{2+}, and Pb^{2+} compounds
Bromides (Br^-)	
Iodides (I^-)	
Sulfates (SO_4^{2-})	Ba^{2+}, Hg_2^{2+}, and Pb^{2+} compounds

"Solubility rules" for Dr Scott's Chem30A class.

Be able to use this table on the exam. A copy will be provided, along with a periodic table.

Similar "rules" may be found in different textbooks etc.

(skip to 5.8) # Net Ionic Equations – Precip. Ex.

Molecular Equation:

$$Pb(NO_3)_2(aq) + 2KI(aq) \rightarrow 2KNO_3(aq) + PbI_2(s)$$

Total Ionic Equation:

$$Pb^{2+}(aq) + \cancel{2NO_3^-(aq)} + \cancel{2K^+(aq)} + 2I^-(aq) \rightarrow$$

$$\cancel{2K^+(aq)} + \cancel{2NO_3^-(aq)} + PbI_2(s)$$

Never break up any (s), (l) or (g) or molecular (aq) species!

Cancel out the spectator ions to yield the net ionic equation:

$$Pb^{2+}(aq) + 2I^-(aq) \rightarrow PbI_2(s)$$

5.5 Neutralization:

Complete ionic: Acid + Base → neutral compounds

Net ionic: $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$

- **Acids** were among the first known "chemicals"
  - Taste sour
  - Turn litmus (extracts of lichens and cabbages) red
  - Evolve a flammable gas (hydrogen) from metals
- **Bases** are "the opposite of acids"
  - Taste bitter
  - Turn litmus solution blue
  - Produce aqueous solutions that feel slippery to the touch
- **Neutral** substances are <u>soluble</u> chemicals (molecules or salts) which are neither acids nor bases
  - Salts are neutral ionic compounds
  - Water (H_2O) is a neutral molecular compound

- The neutralization reaction of an acid with a base yields water plus a *salt*

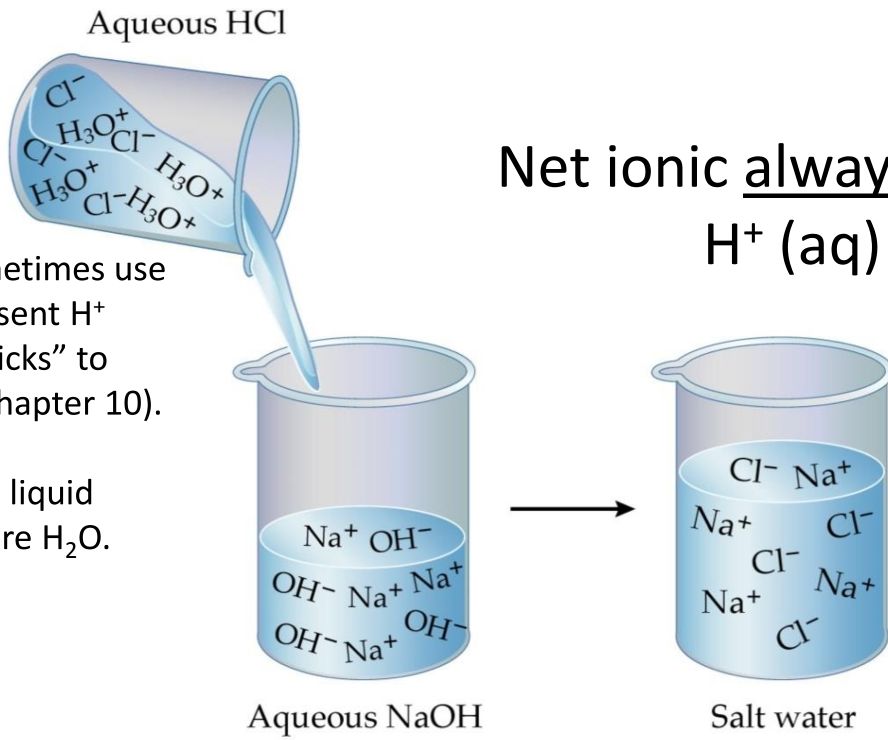
$HCl\ (aq) + NaOH\ (aq) \rightarrow NaCl\ (aq) + H_2O\ (l)$

Net ionic <u>always</u> forms water:

$H^+\ (aq) + OH^-\ (aq) \rightarrow H_2O\ (l)$

Chemists sometimes use H_3O^+ to represent H^+ because it "sticks" to water, H_2O (Chapter 10).

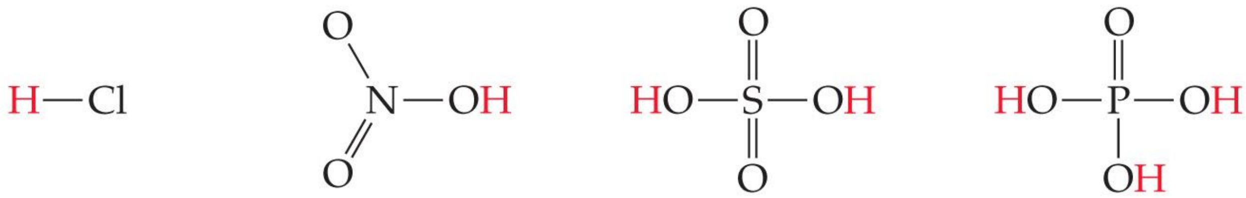
Note the blue liquid represents pure H_2O.



The Na^+ and Cl^- ions do not react… they are <u>spectators</u>. In this case, the salt (NaCl) is soluble and <u>dissolved</u> as ions. That's what NaCl (aq) means – sodium chloride dissolved in water (the aqueous phase).

Common Acids and Bases

- Acids are present in many foods:
  - Lemons, oranges, and grapefruit contain citric acid, and sour milk contains lactic acid
- Bases are present in many household cleaning agents
  - bar soap, ammonia-based window cleaners, drain openers
- Acids can release multiple protons
  - **monoprotic** – single H^+ such as HCl
  - **diprotic** – double H^+ such as H_2SO_4
  - **triprotic** – triple H^+ such as H_3PO_4



Hydrochloric acid (monoprotic) | Nitric acid (monoprotic) | Sulfuric acid (diprotic) | Phosphoric acid (triprotic)

5.7 Oxidation Numbers Rules (for atoms)

1. Each atom in a pure element has an oxidation number of zero
   - Examples of pure elements: Fe (s), Hg (l), Ar (g), O_2 (g)
2. For monatomic ions, the oxidation number is equal to the charge on the ion
   - Example: Na^+ would have an oxidation number of +1
3. For molecules, the oxidation number is similar to charge…
   - Fluorine always has an oxidation number of −1 (except for F_2).
   - The oxidation # of Oxygen is −2
   - Halogens have oxidation number of −1
   - The oxidation # of Hydrogen is (usually) +1…
4. The sum of the oxidation numbers for the atoms equals the charge, or zero for a neutral compound.

5.8 Redox reactions

- A **reducing agent** loses one or more electrons
  - Causes reduction
  - Undergoes oxidation itself
  - Loses electrons to become more positive (or less negative)
- An **oxidizing agent** gains one or more electrons
  - Causes oxidation
  - Undergoes reduction itself
  - **To be <u>reduced</u> means the <u>oxidation number</u> <u>goes down</u>**
  - Gains electrons to become more negative (or less positive)

Redox, for **metals** vs **metals**

+2 electrons = reduced!

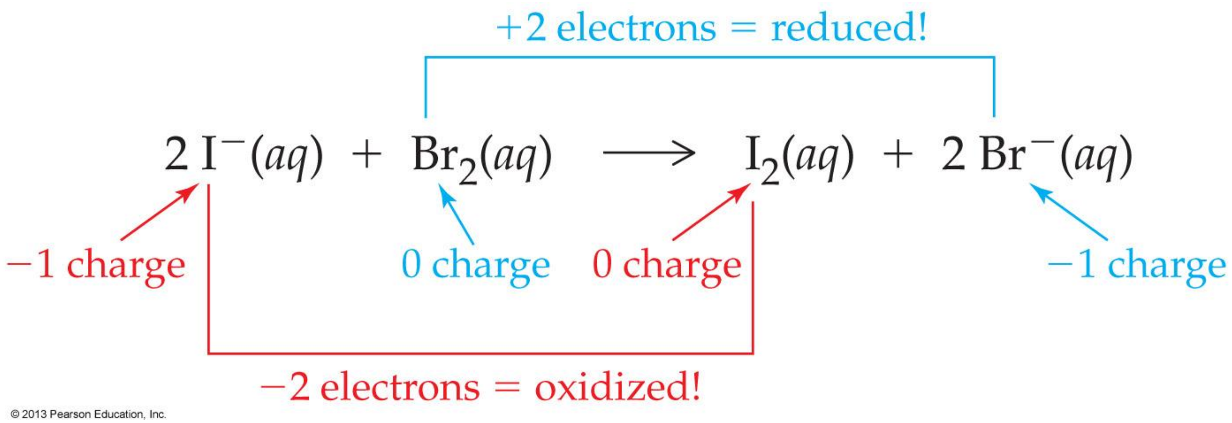
$$Cu(s) + 2\,Ag^{+}(aq) \longrightarrow Cu^{2+}(aq) + 2\,Ag(s)$$

0 charge   +1 charge   +2 charge   0 charge

−2 electrons = oxidized!

This is called plating, not precipitation

Redox, **nonmetals** vs **nonmetals**



- Here, an iodine ion (as in NaI) gives an electron to bromine, forming bromide ions (as in NaBr) and liberating free iodine
  - An iodide ion is oxidized as its charge increases from −1 to 0
  - Bromine is reduced as its charge decreases from 0 to −1

Redox, metals vs nonmetals

- Metals are always oxidized in the presence of a nonmetal
- Nonmetals are always reduced in the presence of a metal

$$Na(s) + Br_2(l) \rightarrow 2\ NaBr(s)$$