Chemistry: What You Need to Know CRSS/FORS 3060

Atoms: All matter is made of atoms; there are different kinds of atoms, called *elements*, which are defined by how many neutrons and protons are in their nucleus. The simplest element is H (hydrogen); it has 1 proton and no neutrons. Electrons orbit the nucleus and balance the charge on the protons; electrons have a -1 charge, protons +1, neutrons are uncharged--neutral (well, yeah . . .). There are 92 naturally occurring elements, each with a unique number of protons (and balancing electrons), and variable numbers of additional neutrons. The number of protons is called the *atomic number*; the sum of the protons and neutrons is called the *atomic mass*. Uranium is the heaviest naturally occurring element, with 92 protons and an average of 146 neutrons (this varies a bit–don't worry about it): its atomic number is 92, its atomic mass is 238.

Protons and neutrons weigh the same; their weight is called an *atomic mass unit*, some very small number. To make weighing elements easier we define a *mole* as 6×10^{23} atoms (called *Avagadro's number*); one mole of hydrogen atoms weighs 1 gram (this is not a coincidence). One mole of uranium atoms weighs 238 grams. This weight is called the *atomic mass (weight)*, in units of g/mole. Convenient, eh?

Molecules: Molecules are combinations of element atoms; they are also called *compounds*. NaCl (sodium chloride) is a compound– common table salt. It forms because Na atom tends to lose one of its electrons to form an *ion*, Na^+; Cl tends to gain an additional electron, forming Cl^-. (Positive ions are called *cations*, negative ones *anions*; the charge on the ions is called the *valence*). These elements like to get together because they satisfy the charge deficit or excess in the other.

We can apply the idea of a mole to compounds as well as atoms; one mole of NaCl contains 1 mole of Na atoms and 1 mole of Cl atoms; the atomic weight of Na is 23, and Cl is 35. Therefore the *molecular weight* (sometimes called the *gram-formula weight)* of NaCl is 58– that is, 58 g per mole (58 g/mole).

Most compounds are solids (altho not necessarily . . .); some are very soluble in water, some less so. Being soluble means the compound breaks up to form ions in the water, forming a *solution*, a single phase containing several components (water, dissolved salt). Adding NaCl to water results in a solution containing Na^+ and Cl^- ions; if 58 g of NaCl were added to 1 liter (L) of water, the resulting solution *concentration* of NaCl would be 1 mole per L (1 mole/L), or a 1 *Molar* solution (abbreviated as *M*).

Reactions: Chemical reactions show how compounds or ions react with each other; there are many types (classes) of reactions, but they all follow basic laws called *stoichiometry*. Consider the reaction: $Na^+ + Cl^-$ _ NaCl As written this reaction shows the *precipitation* of 1 mole of Na and 1 mole of Cl to form one mole of solid compound NaCl. Reactions can go backwards as well, however– note the arrows; going from right to left, the reaction shows the *dissolution* of NaCl to form Na and Cl ions. All reactions are to some degree *reversible* like this. In reactions like this components shows on the left hand side (LHS) are called *reactants*, those on the right hand side (RHS) are *products* of the reaction as written.

You can't easily tell if a reaction is likely to go one way or the other– it depends on a lot of things. Most reactions reach an *equilibrium* point, however, when there is some kind of balance established between reactants and products. You might also say, at equilibrium the *forward* reaction (precipitation of NaCl) is balanced by the *backward* reaction (dissolution of NaCl), such that the composition of the solution stays the same.

Say we take one L of water and add NaCl solid to it; initially there will be *no* Na or Cl ions in the water. The reaction shown above says there is a balance between Na and Cl in solution and solid NaCl in contact with the water. Therefore, the solid will dissolve and the concentration of Na and Cl in solution will increase. This idea is called *Le Chatiler's principle*: adding a *product* to the RHS of an equation will cause the reaction to proceed to the *left*, to try to re-establish this equilibrium. (This also works the other way: adding a reactant to the LHS will force the reaction to move to the *right*). However, if we continue to add solid NaCl, at some point the solid will stop dissolving; the solution will be *saturated* with Na and Cl ions, and no more will dissolve. The solution will then be *in equilibrium*– the composition of the solution will not change.

Charges: Often times in chemical reactions that form or dissolve compounds, it is important to consider the charges on ions in terms of how the reaction proceeds. Consider the reaction of a base, sodium hydroxide (NaOH), with an acid, sulfuric acid (H_2SO_4):

$$2NaOH + H_2SO_4 \text{ --> } 2Na^+ + SO_4^{-2} + 2H_2O$$

Here you need 2 molecules (or moles) of base to provide 2 -OH in order to react with the 2 H^+ atoms (or moles) in each molecule (or mole) of sulfuric acid. The H and OH react to give water (H_2O), with soluble Na and SO_4 remaining in solution.

This "2 moles reacts with 1 mole" thing is unpleasant– it makes doing calculations and keeping track of moles too difficult. So we can define another kind of unit, such as there is a 1:1 stoichiometry clearly shown– in this case, we want 1 H reacting with 1 OH to form 1 H_2O. This unit is called an *equivalent*: it is defined as Avagadro's number (6×10^{23}) of *charges* (rather than atoms or molecules). In an acid/base reaction these charges would be the number of protons (H^+); for sulfuric acid, each mole of H_2SO_4 contains 2 equivalents of H^+. Thus, 1 equivalent of H_2SO_4 can be said to react with 1 equivalent of NaOH . . . . This is most useful when trying to figure out *masses* of reactants and products: say we have 1 mole of NaOH (or, one *equivalent*, since each NaOH has only 1 OH charge); what *mass* of H_2SO_4 will react completely with it? Well, *one equivalent* will; how much does that weigh? The *molecular weight* is (1+1+ 32 + 16+16+16+16=) 98 g/mole; the *equivalent weight* is just this divided by the charge– in this case 2, 98/2=49 g/equiv. Thus 49 g H_2SO_4 will react with one mole of NaOH (which itself, by the way, weighs 23+16+1=40 g/equiv.) Solution concentrations can be expressed in equivalents also; *normality (N)* is defined in equivalents per L, very similar to *M* but using equivalents instead.

The equivalent weight of any ion is just its atomic weight divided by its charge; for compounds, it's the molecular weight divided by the total positive valence. Equivalents are a useful way of thinking about many kinds of reactions– acid/base, redox, and cation exchange, just to name a few.

Table of Atomic Masses (you don't need to memorize these.....they will be given to you).

H (hydrogen)	1	P(phosphorus)	31	Al^{+3}	27
O(oxygen)	16	Cl(chlorine)	35	Ca^{+2}	40
C(carbon)	12	N(nitrogen)	14	K^+	39
S(sulfur)	32			Na+	23
				Mg^{+2}	24

Concentration units: mg/L = µg/mL = ppm (solution) mg/kg = µg/g = ppm (solid [soil])
meq/100g = cmol(±)/kg = units of CEC [meq
M = moles/L = millimoles/mL [molarity]
N = equivalents/L = milliequivalents/mL [normality]

Some Problems . . .

1. Describe how to make up 1 L of: A. 0.5 *M* KNO_3 B. 2 *M* $CaCl_2$ C. 0.005 *M* NaOH
2. Calculate the equivalent masses (g/equivalent) for: A. Ca^{+2} B. Na_2SO_4 C. $Al(OH)_3$
3. *Titrations:* remember that at the endpoint (neutral pH), equivalents of acid (or base) added must equal equivalents of base (or acid) initially present in solution.
   A. How many equivalents of acid are in a solution which requires 12 mL of 0.005 *N* NaOH to neutralize?
   B. What is the concentration of base (*N*) if 50 mL of a base solution is titrated to the endpoint using 22 mL of 0.06 *N* H_2SO_4?
   C. Acidic cations are extracted from 10 g of soil using 50 mL of salt solution; 35 mL of this solution is titrated to the endpoint using 3.8 mL of 0.008 *N* NaOH. Calculate the exchangeable acidity in this soil in milliequivalents/100 g soil.
4. *Exchangeable Cations:* 10 g of soil is extracted with 50 mL of a salt solution to displace exchangeable cations; Mg is measured by AA and found to be 4 ppm in the extracting solution. Calculate how much exchangeable Mg was present in the soil in meq/100 g. Al was also measured in this solution, and found to be 12 ppm. Calculate exchangeable Al. If Ca= 2.2 meq/100 g and K=0.12 meq/100 g, calculate CEC and %BS [assuming no exchangeable H].
5. Calculate the following:
   A. 30 ppm KNO_3 (in solution) is how many meq/L?
   B. A soil has 0.5 meq K/100 g; how many lbs. K/afs is this? [assume afs=1,000 tons]
   C. A waste product has 50 ppm Cd in it; if you apply 10 tons of this per acre, how much Cd have you added to the soil in ppm?