

Prototypical Reactions

Chemistry 35
Fall 2000

Solution Concentrations

- We need to be able to *quantify* amounts of compounds in solutions:

1. **Mass Percent**

$$\text{Mass Percent} = \frac{\text{Solute Mass}}{\text{Solution Mass}} \times 100$$

Used, more typically, for very dilute solutions:

$$\text{ppm} = \frac{\text{Solute Mass}}{\text{Solution Mass}} \times 10^6 \quad \text{Trace}$$

Concentrations: Moles

- Since reaction chemistry is quantified using moles, these are more useful:

2. **Mole Fraction** (X) = $\frac{\text{mol solute}}{\text{Total mol}}$

3. **Molarity** (M) = $\frac{\text{mol solute}}{\text{Liters Solution}}$

4. **Molality** (m) = $\frac{\text{mol solute}}{\text{kg solvent}}$

-temp independent

3

Molarity

- Typical concentrations:

-usually won't see much below about 10^{-2} M

What is the **molarity** of conc. H_2SO_4 ?

on the bottle: 94.0 % H_2SO_4 - mass %

1.831 g/mL - density

A conversion: mass % @ molarity

$$\begin{aligned} \frac{94.0 \text{ g H}_2\text{SO}_4}{100.0 \text{ g sol'n}} \times \frac{1 \text{ mol H}_2\text{SO}_4}{98.08 \text{ g H}_2\text{SO}_4} \times \frac{1.831 \text{ g sol'n}}{\text{mL sol'n}} \times \frac{1000 \text{ mL sol'n}}{1 \text{ L sol'n}} &= \\ &= 17.54833 \text{ mol H}_2\text{SO}_4/\text{L sol'n} \\ &= \underline{\underline{17.5 \text{ M H}_2\text{SO}_4}} \end{aligned}$$

4

Molarity of Water

- In aqueous solutions, the highest concentration is the molarity of pure water:

How many mol H₂O per Liter?

$$\begin{aligned} 1.000 \text{ L H}_2\text{O} \times \frac{1000 \text{ mL}}{1 \text{ L}} \times \frac{1.00 \text{ g H}_2\text{O}}{1 \text{ mL}} \times \frac{1 \text{ mol H}_2\text{O}}{18.0148 \text{ g H}_2\text{O}} &= \\ &= 55.50991 \text{ mol H}_2\text{O} \end{aligned}$$

$$\text{So: } \frac{55.5 \text{ mol H}_2\text{O}}{1.000 \text{ L H}_2\text{O}} = \underline{\underline{55.5 \text{ M H}_2\text{O}}}$$

5

Dilutions

- # mol solute remains unchanged during dilution:

$$\# \text{ mol solute (initial)} = \# \text{ mol solute (final)}$$

concentrated

diluted

$$M_1 V_1 = M_2 V_2$$

Example: How many mL conc. H₂SO₄ must be diluted to make 500. mL of 2.0 M H₂SO₄?

$$17.54844 \text{ M (V}_1) = 2.0 \text{ M (500. mL)}$$

$$V_1 = 56.985 \text{ mL}$$

-Add 57. mL conc. H₂SO₄ to water and bring up to 500. mL

6

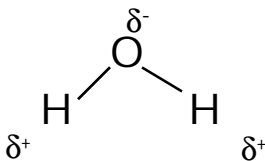
Reactions in Aqueous Solution

- Reactions in **water** are especially of interest (for obvious reasons), so let's look at:
 - Properties of Aqueous Solutions
 - Acid/Base Reactions in Water
 - Solubility of Compounds in Water
 - Oxidation/Reduction (Redox) Reactions

7

Properties of Aqueous Solns

- Water is a **Polar** molecule:



-*negatively* charged oxygen will attract positive ions in solution

-*positively* charged hydrogens will attract negative ions in solution

8

Electrolytes

- **Ionic compounds** break up into *ions* in aqueous solution and are called electrolytes.

Examples:

- **NaCl** - *strong* electrolyte (dissolves completely)
 - **HCl** - *strong* electrolyte (dissolves completely)
 - **HC₂H₃O₂** - *weak* electrolyte (partially dissolves)
- **Molecular compounds** do not ionize when they dissolve in water and are called nonelectrolytes.

9

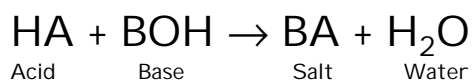
Acid/Base Reactions

- Most common definition:

Bronsted-Lowry

Acid: *proton donor*

Base: *proton acceptor*



-an example of a *double displacement reaction (metathesis)*

10

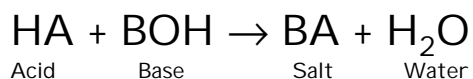
Acid/Base Reactions

- Most common definition:

Bronsted-Lowry

Acid: *proton donor*

Base: *proton acceptor*



-an example of a *double displacement reaction (metathesis)*

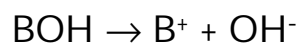
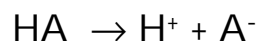
11

Bronsted-Lowry in Water: Arrhenius Acids and Bases

- In aqueous solution, we can use the Arrhenius definitions of acids and bases:

Acid: *substance yielding H^+ in water*

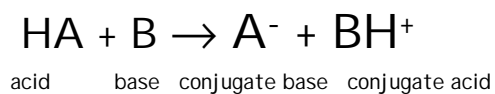
Base: *substance yielding OH^- in water*



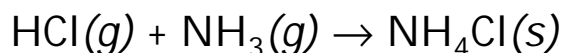
12

More B-L Acid/Base

- Bronsted-Lowry does not require reaction to be in aqueous solution:



Example:



13

Lewis Acids and Bases

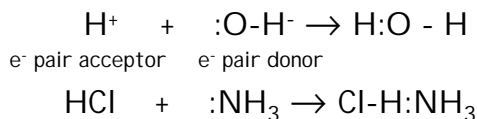
- We can make this concept *broader* by defining acids and bases a bit differently:

Lewis Theory

Acid: electron pair *acceptor*

Base: electron pair *donor*

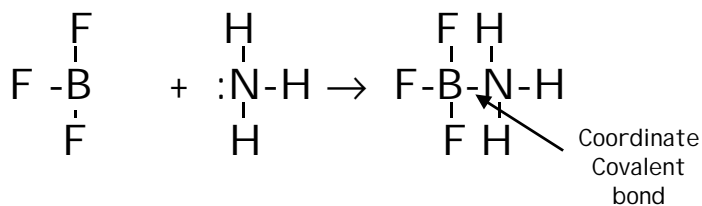
Fits with Bronsted-Lowry definition:



14

More Lewis Theory

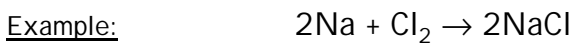
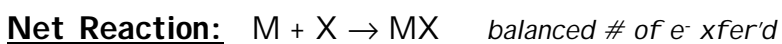
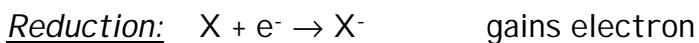
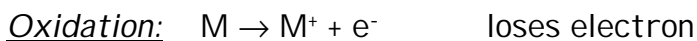
- *MORE* reactions can be understood as acid/base reactions:



15

Oxidation/Reduction Reactions

- Reactions in which electrons are transferred *individually* (not as pairs):



-Na is *oxidized* to Na^+ (gives up an e^- to Cl)

-Cl is *reduced* to Cl^- (takes e^- from Na)

16

Oxidation States

- We can assign a charge to atoms in a compound based on their propensity for donating or accepting valence electrons:

OXIDATION NUMBER

Examples: NaCl - Na \rightarrow +1; Cl \rightarrow -1
K₂O - K \rightarrow +1; O \rightarrow -2

- these assignments do not necessarily reflect the *actual distribution of electrons* in a compound (they are just a way to keep track of the valence electrons)

17

Assigning Oxidation Numbers

- **Some rules:**
 - Sum of oxidation # = overall charge
 - **Group I, II:** always +1, +2 (ionic cmpds)
 - **Group VII:** F *always* -1, others often -1; could be + in cmpds with O or other Group VII elements
 - **Hydrogen:** always +1, except when with Group I, II (then it's -1; hydride)
 - **Oxygen:** always -2, except with F, hydrogen or Group I, II, or with itself (see exceptions in book)

18

Oxidation #: examples

- SrBr_2 : group II, group VII

+2 -1

- Zn(OH)_4^{2-} : OH^- is -1, -2 charge for compd

+2 -2 +1

- $\text{Cr}_2\text{O}_7^{2-}$: $-2 = 7(-2) + 2(x)$

+6 -2 $x = +6$

- MnO_4^- : $-1 = 4(-2) + x$

+7 -2 $x = +7$

19

Formal Charges

- Another way of looking at charge distribution in compounds:

-whereas *oxidation numbers* assume complete transfer electrons between atoms (ionic model)

-*formal charges* assumes complete **sharing** of electrons between atoms (covalent model)

Formal Charge =

$\# \text{valence } e^- - (\# \text{bonds} + \# \text{nonbonding } e^-)$

20

Formal Charges: Example

- What is the structure of HCN?

Two choices:



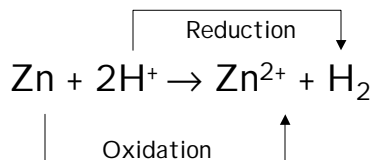
$$\begin{array}{ll} 1) \underline{\text{H}}: 1-(1+0) = 0 & 2) \underline{\text{H}}: 1-(1+0) = 0 \\ \underline{\text{C}}: 4-(4+0) = 0 & \underline{\text{C}}: 4-(3+2) = -1 \\ \underline{\text{N}}: 5-(3+2) = 0 & \underline{\text{N}}: 5-(4+0) = +1 \end{array}$$

Structure which minimizes formal charges is preferred

21

Back to Redox Reactions

- Redox reactions *change* the oxidation states of the reactants:



Zn: is *oxidized* (reducing agent)

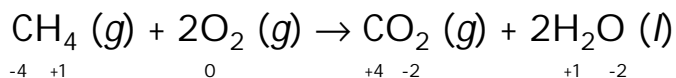
H⁺: is *reduced* (oxidizing agent)

Reox movies

22

Redox Examples

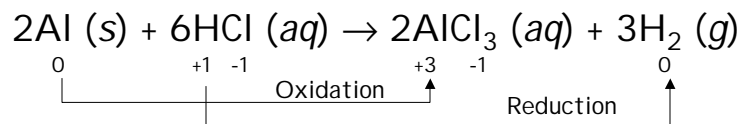
■ Combustion



C: oxidized (-4 → +4)

O: reduced (0 → -2)

■ Acid oxidation of a metal



23

Solubility

- How do we know whether a compound is soluble in water?

-we will address this from a *structural* point of view in Chem 36

NaCl dissolution movie

-for now, we need to learn **Solubility Rules**:

24

Solubility Rules: Goldberg

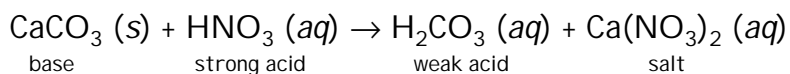
■ ALL are Soluble:

- **Nitrates** (NO_3^-)
- **Chlorates** (ClO_3^-)
- **Acetates** (CH_3COO^-)
- **Sulfates** (SO_4^{2-})
 - **Except:** BaSO_4 , SrSO_4 , PbSO_4 <- *insoluble*
 CaSO_4 , Ag_2SO_4 , Hg_2SO_4 <- *slightly insoluble*
- **Chlorides, Bromides and Iodides**
 - **Except:** Ag^+ and Hg^+ salts <- *insoluble*
 Pb^{2+} Salts <- *slightly insoluble*

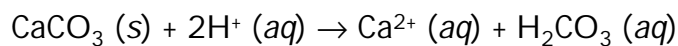
25

Acid/Base? Solubility? Both!

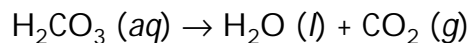
■ Calcium Carbonate + Nitric Acid



Net Ionic:



But, also:



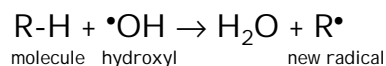
26

Radicals Beget Radicals

- **Radical:** chemical species containing an *unpaired* electron

Example: atomic Cl - 7 electrons (1 unpaired)

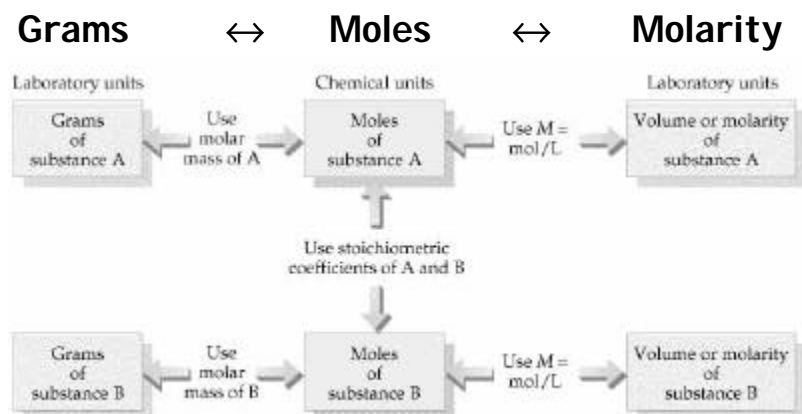
-Radicals are VERY reactive and will aggressively seek out another atom to pair up the electron and form a covalent bond:



-leads to: *chain reactions*

27

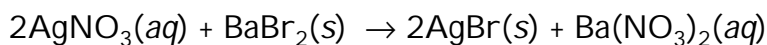
Solution Stoichiometry



28

Example Problems

- What *volume* of 0.100 M AgNO₃ will precipitate all of the Br in a 2.000-g sample of BaBr₂?



grams BaBr₂ → mol BaBr₂ → mol AgNO₃ → L AgNO₃ → mL AgNO₃

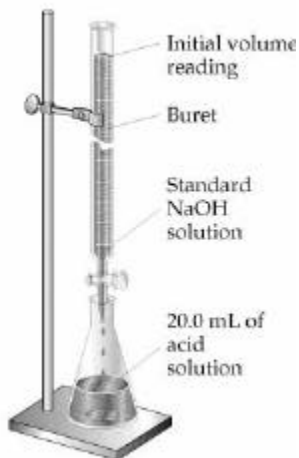
$$\begin{aligned} 2.000 \text{ g BaBr}_2 \times \frac{1 \text{ mol BaBr}_2}{297.16 \text{ g BaBr}_2} \times \frac{2 \text{ mol AgNO}_3}{1 \text{ mol BaBr}_2} \times \frac{1 \text{ L AgNO}_3}{0.100 \text{ mol AgNO}_3} \times \frac{1000 \text{ mL}}{1 \text{ L}} &= \\ &= 1.346076 \times 10^2 \text{ mL AgNO}_3 \text{ soln} \\ &= \underline{\underline{1.35 \times 10^2 \text{ mL AgNO}_3 \text{ soln}}} \end{aligned}$$

29

Titration

- Quantitative reactions performed so as to reach *stoichiometric equivalence* at a determinable endpoint

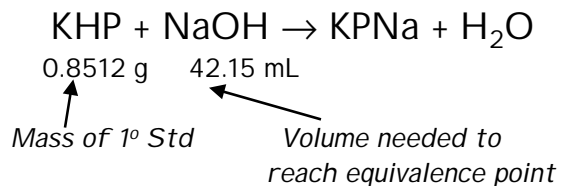
HOW?



30

Titration: Example

1. Standardization of NaOH solution



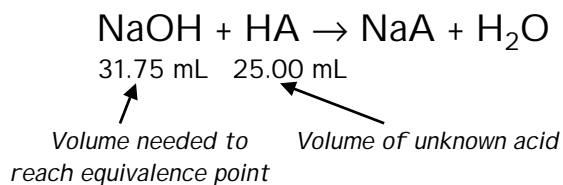
What's the concentration of the NaOH solution?

$$\frac{0.8512 \text{ g KHP}}{0.04215 \text{ L NaOH}} \times \frac{1 \text{ mol KHP}}{204.23 \text{ g KHP}} \times \frac{1 \text{ mol NaOH}}{1 \text{ mol KHP}} = \mathbf{0.09888 \text{ M NaOH}}$$

31

Titration: Example (cont'd)

2. Titration of Sample



What's the concentration of the acid solution?

$$\frac{31.75 \text{ mL NaOH}}{25.00 \text{ mL HA}} \times \frac{0.0988814 \text{ mol NaOH}}{\text{L}} \times \frac{1 \text{ mol HA}}{1 \text{ mol NaOH}} = \mathbf{0.1256 \text{ M HA}}$$

32