7 • Chemical Reactions & the Per. Table
Solutions and Solubility
(1 of 16)

We learned the terms solute, solvent, and solution.
Solubility (how MUCH solute will dissolve) is measured in g/100 mL of H_2O or sometimes in M.
This information may be given numerically or graphically.
A **saturated** solution is one in which any additional solute added will simply settle on the bottom of the container.
An **unsaturated** solution is any amount less than saturated.
Supersaturated implies that the solution was saturated at some **higher** temperature and then carefully **cooled**. This **unstable** situation can be changed with a "**seed**" crystal.
Recall the supersaturated solution of $NaC_2H_3O_2$ demo.
The terms **concentrated** and **dilute** refer to the amount of solute and do not necessarily coincide with saturation.

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7 • Chemical Reactions & the Per. Table
Electrolytes: Weak and Strong
(2 of 16)

Solutions of **acids**, **bases**, and **salts** contain **mobile ions** and conduct electricity. These solutions are called **electrolytes**.
Salts are **ionic** compounds that **dissociate** in water.
Acids are actually **molecular** compounds (covalently bonded) the become ions when dissolved in water.
Only 8 **acids** are **strong** electrolytes and **completely dissociate** when dissolved. All others dissolve completely, but only partially dissociate into ions.
Only 8 **bases** are **strong** electrolytes because they **dissolve completely**. All others have low solubility and remain **solids** rather than dissolve. One common **exception** is the weak base **NH_4OH** . It **dissolves**, but partially dissociates.

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7 • Chemical Reactions & the Per. Table
Ionic Reactions
(3 of 16)

A common class of chemical reactions occurs when two **ionic** solutions are mixed. The **double replacement** or **metathesis** reaction involves the formation of two new combinations of ions.
$AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$ (molecular equation)
The **new** combinations may be **more stable** than the original due to **low solubility** (**precipitate** forms), **weak electrolyte**, **gas** formation, or **complex ion** formation.
The reaction is written above as though the substances exist as **molecules**. This is the easiest time to balance.
The **ionic equation** shows **strong** electrolytes as **separate ions.** The **net equation** eliminates "spectator ions".

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7 • Chemical Reactions & the Per. Table
Arrhenius Acids and Bases
(4 of 16)

Acid: a substance that increases the **[H+]** in solution.
Base: a substance that increases the **[OH–]** in solution.
Diprotic acids have more than one removable H. (H_2SO_4)
However, only the first H is ever easily dissociated.
Oxides of **nonmetals** (SO_2) are **acid anhydrides**.
Oxides of **metals** (Na_2O)are **basic anhydrides**.
Just add water... to get the acid or base.
$SO_2 + H_2O \rightarrow H_2SO_3$ $Na_2O + H_2O \rightarrow 2\ NaOH$
Acids and bases **neutralize** each other because the H+ and OH– ions form the very weak electrolyte... H_2O (and a salt).
Acid salts are the partially neutralized polyprotic acids.
NaH_2PO_4 or Na_2HPO_4 or $NaHSO_4$, etc... solid acids.

7 • Chemical Reactions & the Per. Table
Brønsted-Lowry Acids and Bases
(5 of 16)

Acid: a proton (H^+) **donor**. **Base**: a proton **acceptor**.
This is a more general definition of acids and bases because it does not **require** the substance to be dissolved in water.
Consider the following equations:
The species that accepted the proton (base) can be considered a donor (conjugate acid) in the <u>reverse</u> reaction.

$$\underset{\text{(acid)}}{HF} + \underset{\text{(base)}}{H_2O} \Leftrightarrow \underset{\text{(acid)}}{H_3O^+} + \underset{\text{(base)}}{F^-}$$

$$\underset{\text{(base)}}{NH_3} + \underset{\text{(acid)}}{H_2O} \Leftrightarrow \underset{\text{(acid)}}{NH_4^+} + \underset{\text{(base)}}{OH^-}$$

Strong base≈weak conjugate acid, (good acceptor≈lousy donor). Conjugates differ by only a H^+ (e.g. HF and F^-)

7 • Chemical Reactions & the Per. Table
Ions in Water
Some Metal Ions Make Acidic Solutions
(6 of 16)

Since water molecules are polar, they surround ions in solution (called **hydration**). When we write $Na^+(aq)$ and $Cl^-(aq)$ we are implying this more complex situation.
Some highly charged ions (Al^{3+}, Cr^{3+}, Fe^{3+}) tend to tightly bind the water molecules. We can write them as complex ions: $Al(H_2O)_6^{3+}$
The electron clouds are drawn toward the central ion and away from the oxygen and therefore the O-H bond. This extra-polar O-H bond results in the H atom more readily joining with a passing water molecule... making the solution acidic. [Note: this is a conjugate acid-base situation.]

$$Al(H_2O)_6^{3+} + H_2O \Leftrightarrow H_3O^+ + Al(H_2O)_5OH^{2+}$$

7 • Chemical Reactions & the Per. Table
Trend in Strengths of Acids and Bases
Three cases to explain
(7 of 16)

Case 1: The more oxygens on an oxoacid, the stronger the acid. $H_2SO_4 > H_2SO_3$.. $HClO_4 > HClO_3 > HClO_2 > HClO$
Why?: The electronegative oxygens draw electron density away from the central atom and therefore from the H-O bond... making it more polar. H leaves more easily.
Case 2: The more electronegative the central atom, the stronger the oxoacid. $H_3PO_4 < H_2SO_4 < HClO_4$

Why? Same as above... the more electronegative atom in the center makes the O-H bond more polar.
Case 3: Binary acid strength depends on the **SIZE** of the atom... HF < HCl < HBr < HI... not the electronegativity.
Why?: The greater distance means a weaker attraction.

7 • Chemical Reactions & the Per. Table
Lewis Acids and Bases
(8 of 16)

Consider the formation of a **coordinate covalent bond**:
Acid: electron pair acceptor [Note: proton *donor*]
Base: electron pair donor [Note: proton *acceptor*]
This definition is more general than the other two because this covers cases that don’t even involve hydrogen (protons).
Classic case: $HN_3 + BF_3 \rightarrow NH_3BF_3$
Other important cases:

oxide of metal + oxide of nonmetal → salt
(base anhydride + acid anhydride)

$$CaO + SO_2 \rightarrow CaSO_3(s)$$

and the oft-confusing reactivity of CO_2

$$OH^- + CO_2 \rightarrow HCO_3^-$$

7 • Chemical Reactions & the Per. Table
Oxidation Numbers
(9 of 16)

Definition:
Oxidation numbers are the *apparent* charges atoms have if shared e^-'s are assigned to the *more* electronegative atom. You can assign ox. #'s by studying a Lewis diagram or...

The Rules:
…ox. # of neutral atoms is 0
In compounds…
…simple ions have ox. #'s equal to their charge.
…F (-1), Family I (+1), II (+2), Al (+3)
…O is usually (-2) except peroxide and OF_2
…H is usually (+1) except hydrides of Fam. I, II, Al
…the sum of the ox. #'s of individual atoms equals the charge on the entire species.

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7 • Chemical Reactions & the Per. Table
Balancing Redox Equations
Oxidation Number Change Method
Step-By-Step
(10 of 16)

1. Identify the ox. #'s of elements that change PER ATOM. One element will change UP (oxidation / lose e^-'s) one element changes DOWN (reduction / gain e^-'s).
2. Adjust coefficients and e^- changes for situations where more than one atom MUST change:
   ex: $\mathbf{2}H\mathbf{Cl} + K_2\mathbf{Cr}_2O_7 \rightarrow KCl + \mathbf{2Cr}Cl_3 + \mathbf{Cl}_2 + H_2O$
   Cr: 2 x (3 e^- per Cr)   Cl: 2 x (1 e^- per Cl)
3. Balance these changes in e^-'s (e^- gained = e^- lost).
4. Balance all elements except H and O.
5. Balance O's (add H_2O's as needed).
6. Balance H's (add H^+'s as needed).
7. If solution is basic, see card #12.

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7 • Chemical Reactions & the Per. Table
Balancing Redox Equations
Ion-Electron Method
(Half-reaction Method)
Step-By-Step
(11 of 16)

1. Identify the substances involved in the oxidation and reduction changes. Include substances so that all elements are represented (except H and O).

For each half-reaction…

2. Balance all elements except H and O.
3. Balance O's (add H_2O's as needed).
4. Balance H's (add H^+'s as needed).
5. Balance charges (add e^-'s to the more positive side).
6. Balance e^-'s in the two half-reactions.
7. Combine the two half-reactions. Cancel substances that show up on both sides of the equation (e^-'s must cancel).
8. If solution is basic, see card #12.

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7 • Chemical Reactions & the Per. Table
Balancing Redox Equations
Reactions in Basic Solution
(12 of 16)

If the reaction occurs in a basic solution (usually stated clearly in the problem) then instead of H^+'s and H_2O's, you utilize OH^-'s and H_2O's.

An easy method is:
- balance as though the reaction were in an acid solution.
- add OH^-'s to each side of the equation until all H^+'s are turned into H_2O's.
- cancel H_2O's that show on both sides of the equation.

7 • Chemical Reactions & the Per. Table
Metals as Reducing Agents
(13 of 16)

In each case, the metal changes to the + ion: $M \rightarrow M^+ + e^-$
Since the metal is oxidized, it is a reducing agent.
Metals that most easily lose e^-'s (those with low ionization energy and low electronegativity) make the best reducers.

MEMORIZE THIS:

Some metals react with H_2O [to $H_2(g)$ and OH^-]
Some react with non-oxidizing acids such as HCl, and cold, dilute H_2SO_4 (the H^+ is the reacting species) [to $H_2(g)$]
Some react only with oxidizing acids:

- dilute HNO_3 [to colorless $NO(g) + H_2O$],
- conc. HNO_3 [to red-brown $NO_2(g)$ and H_2O], and
- hot, conc. H_2SO_4 [to $SO_2(g)$ and H_2O]

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7 • Chemical Reactions & the Per. Table
The Activity Series of Metals
(14 of 16)

FOUR GROUPS OF METALS

1 - Most active - Families I and II - great reducing agents
  reduction half-reaction: $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$
2 - Most metals... Zn, Fe, Al, etc.
  reduction half-reaction: $2H^+ + 2e^- \rightarrow H_2$
3 - Ag, Cu, Hg
  ex. half-reaction: $2H^+ + NO_3^- + e^- \rightarrow NO_2 + H_2O$
4 - Noble Metals - Au, Pt, Ir
  only changed by "aqua regia" [HNO_3 + HCl forms Cl_2]

A more active metal can reduce or displace the ion of a less active metal. Ex. Zn° is more active than Cu°, so…

$Zn^\circ + Cu^{2+} \rightarrow Zn^{2+} + Cu^\circ$ $Cu^\circ + Zn^{2+} \rightarrow$ no reaction

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7 • Chemical Reactions & the Per. Table
Nonmetals as Oxidizing Agents
Oxygen as an Oxidizing Agent
Combustion as a Redox Reaction
(15 of 16)

Elemental **nonmetals** (O, Cl, F, S, etc.) form negative ions by gaining electrons (reduction) and are oxidizing agents.

The strongest oxidizing agents are those to the right of each period (excluding noble gases) and those at the top of each family. So we can predict that $F > Cl > Br > I$ and $O > S$.

Example: Cl will displace Br^-, but not F^-
$Cl_2(g) + 2NaBr \rightarrow Br_2(l) + 2NaCl$ • $Cl_2 + NaF \rightarrow$ N.R.

Oxygen, O_2, is a common and powerful oxidizing agent. Corrosion of metals, formation of oxides, and combustion are all examples of redox with O_2 as the oxidizer.

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7 • Chemical Reactions & the Per. Table
Amphiprotic/Amphoteric & Leveling Effect
(16 of 16)

Anything with a **lone pair** can act as a proton **acceptor** (base) Anything with a **H atom** can act as a proton **donor** (acid). See **water** on card 5 act as either an acid or a base. See acetic acid on page 241 act as both acid and base. We say these are **amphiprotic** or **amphoteric** substances.

Leveling effect...

You cannot tell which strong acid is <u>strongest</u> in water, because donating a proton to water is not a good enough challenge. Strong acids completely dissociate in water because water is a pretty good acceptor of protons. We say that water has a **leveling effect.** Acetic acid is amphiprotic, but a poor proton acceptor and therefore is a great test for which strong acid is the strongest...