

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

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We learned the terms solute, solvent, and solution.  
**Solubility** (how MUCH solute will dissolve) is measured in g/100 mL of H<sub>2</sub>O 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<sub>2</sub>H<sub>3</sub>O<sub>2</sub> 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  
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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<sub>4</sub>OH**. It **dissolves**, but partially dissociates.

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7 • Chemical Reactions & the Per. Table  
Ionic Reactions  
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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.  
 $\text{AgNO}_3 + \text{NaCl} \rightarrow \text{AgCl} + \text{NaNO}_3$  (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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Arrhenius Acids and Bases  
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**Acid**: a substance that increases the [H<sup>+</sup>] in solution.  
**Base**: a substance that increases the [OH<sup>-</sup>] in solution.  
**Diprotic** acids have more than one removable H. (H<sub>2</sub>SO<sub>4</sub>)  
*However, only the first H is ever easily dissociated.*  
**Oxides of nonmetals** (SO<sub>2</sub>) are **acid anhydrides**.  
**Oxides of metals** (Na<sub>2</sub>O) are **basic anhydrides**.  
*Just add water... to get the acid or base.*  
 $\text{SO}_2 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{SO}_3$     $\text{Na}_2\text{O} + \text{H}_2\text{O} \rightarrow 2 \text{NaOH}$   
Acids and bases **neutralize** each other because the H<sup>+</sup> and OH<sup>-</sup> ions form the very weak electrolyte... H<sub>2</sub>O (and a salt).  
**Acid salts** are the partially neutralized polyprotic acids.  
NaH<sub>2</sub>PO<sub>4</sub> or Na<sub>2</sub>HPO<sub>4</sub> or NaHSO<sub>4</sub>, etc... solid acids.

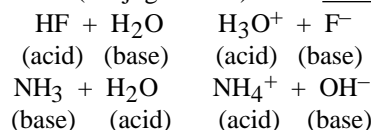
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Brønsted-Lowry Acids and Bases  
(5 of 16)**

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**Acid:** a proton (H<sup>+</sup>) **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 reverse reaction.



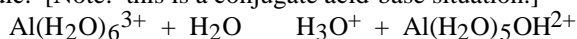
Strong base weak conjugate acid, (good acceptor lousy donor). Conjugates differ by only a H<sup>+</sup> (e.g. HF and F<sup>-</sup>)

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**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<sup>+</sup>(aq) and Cl<sup>-</sup>(aq) we are implying this more complex situation. Some highly charged ions (Al<sup>3+</sup>, Cr<sup>3+</sup>, Fe<sup>3+</sup>) tend to tightly bind the water molecules. We can write them as complex ions: Al(H<sub>2</sub>O)<sub>6</sub><sup>3+</sup>

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.]



**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<sub>2</sub>SO<sub>4</sub> > H<sub>2</sub>SO<sub>3</sub>.. HClO<sub>4</sub> > HClO<sub>3</sub> > HClO<sub>2</sub> > 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<sub>3</sub>PO<sub>4</sub> < H<sub>2</sub>SO<sub>4</sub> < HClO<sub>4</sub>

**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.

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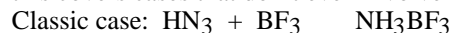
**7 • Chemical Reactions & the Per. Table  
Lewis Acids and Bases  
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Consider the formation of a **coordinate covalent bond**:

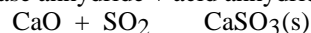
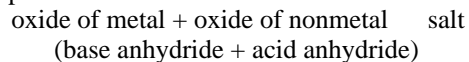
**Acid:** electron pair acceptor [Note: proton *donor*]

**Base:** electron pair donor [Note: proton *acceptor*]

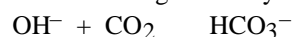
This definition is more general than the other two because this covers cases that don't even involve hydrogen (protons).



Other important cases:



and the oft-confusing reactivity of CO<sub>2</sub>



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**7 • Chemical Reactions & the Per. Table  
Oxidation Numbers  
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**7 • Chemical Reactions & the Per. Table  
Balancing Redox Equations  
Oxidation Number Change Method  
Step-By-Step  
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**7 • Chemical Reactions & the Per. Table  
Balancing Redox Equations  
Ion-Electron Method  
(Half-reaction Method)  
Step-By-Step  
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**7 • Chemical Reactions & the Per. Table  
Balancing Redox Equations  
Reactions in Basic Solution  
(12 of 16)**

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**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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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:  $2HCl + K_2Cr_2O_7 \rightarrow KCl + 2CrCl_3 + Cl_2 + H_2O$   
Cr:  $2 \times (3 e^- \text{ per Cr})$     Cl:  $2 \times (1 e^- \text{ per Cl})$
  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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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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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.
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**7 • Chemical Reactions & the Per. Table  
Metals as Reducing Agents  
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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  
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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.  $Zn^\circ$  is more active than  $Cu^\circ$ , 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  
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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 strongest 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...

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