Acids, Bases and Alkalis Cheat Sheet

by fongrsy via cheatography.com/65383/cs/16397/

Definitions

Acids	Acids are compounds which ionise/dissociate in water to produce hydrogen ions (H^+).
Bases	Bases are compounds that are metal oxides or hydroxides that react with an acid to give a salt and water only.
Alkalis	Alkalis are bases that ionise/dissociate in water to produce hydroxide ions (OH^-).

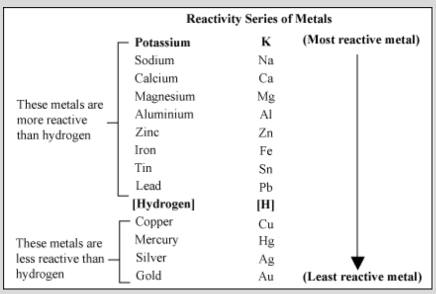
Examples of Acids & Bases

Acid	Chemical Formula	Base	Chemical Formula
Hydrochloric Acid	HCl	Magnesium Oxide	MgO
Sulfuric Acid	H_2SO_4	Copper (II) Oxide	CuO
Nitric Acid	HNO_3	Sodium Hydroxide	$NaOH$
Citric Acid	$C_6H_8O_7$	Potassium Hydroxide	KOH
Ethanoic Acid	CH_3CO_2H	Calcium Hydroxide	$Ca(OH)_2$
Lactic Acid	$C_3H_6O_3$	Aqueous Ammonia	NH_3

Acids 1 to 3 are known as mineral / inorganic acids while Acids 4 to 6 are known as organic acids.

Bases 1 & 2 are insoluble bases while Bases 3 to 6 are soluble bases / alkalis.

Metal Reactivity Series



Types of Reactions

Metal + Acid → Salt + Hydrogen Gas
Metal Carbonate + Acid → Salt + Water + Carbon Dioxide
Metal Oxide + Acid → Salt + Water
Metal Hydroxide + Acid → Salt + Water
Base + Acid → Salt + Water (Neutralisation)
Alkali + Acid → Salt + Water (Neutralisation)
Alkali + Ammonium Salt → Salt + Water + Ammonia Gas
Alkali + Salt → Metal Hydroxide + Salt

Tests for Gases:
Hydrogen Gas - Extinguishes a lighted splinter with a 'pop' sound.
Carbon Dioxide Gas - Released as effervescence. Reacts with limewater to form a white precipitate.
Ammonia Gas - Pungent odour. Turns red litmus paper blue.

Notes:
Base / Alkali + Acid is an exothermic reaction.
Pb (s) + H_2SO_4 / HCl → $PbSO_4$ / $PbCl_2$ + H_2
Lead reacts slowly then stops. Salt forms on the surface of the lead. The salt formed is insoluble.

pH Scale

Acidic solutions have pH values < 7.
They contain **more** H^+ ions and **fewer** OH^- ions.
Neutral solutions have pH values = 7.
They contain **equal amounts** of H^+ ions and OH^- ions.
Alkaline solutions have pH values > 7.
They contain **more** OH^- ions and **fewer** H^+ ions.

Ionic Equations

1. Write a balanced chemical equation with state symbols.
2. Check which reactants and products can form ions in water. (Aqueous)
3. Split up these reactants and products into their respective ions.
4. Check for ions that appear in both LHS & RHS of the equation, these are spectator ions that can be removed from the equation.

Ionic Equations (cont)

5. For those reactants and products which are unable to form ions, do not split the compounds.
6. What is left will be the net ionic equation. The coefficients must be in the lowest ratio.

Polyatomic Ions

Charge	Name	Chemical Formula
1+	Ammonium	NH_4^+
	Hydronium	H_3O^+
1-	Nitrate	NO_3^-
	Hydroxide	OH^-
	Ethanoate	CH_3COO^-
2-	Carbonate	CO_3^{2-}
	Sulfate	SO_4^{2-}
3-	Phosphate	PO_4^{3-}

Notes:
Silver ion: Ag^+
Zinc ion: Zn^{2+}

Properties of Acids

1. Acids are corrosive.
2. Acids have a sour taste.
3. Acidic solutions conduct electricity. (Electrolytes)
4. Acids change the colour of indicators.
Litmus Paper: Blue to Red
Methyl Orange Solution: Orange to Red
Universal Indicator Paper: Orange to Red
Universal Indicator Solution: Green to Red

Properties of Alkalis

1. Alkalis have a soapy feeling and a bitter taste.
2. Alkaline solutions conduct electricity. (Electrolytes)
3. Alkalis change the colour of indicators.
Litmus Paper: Red to Blue
Methyl Orange Solution: Orange to Yellow
Universal Indicator Paper: Orange to Violet
Universal Indicator Solution: Green to Violet

Balancing Chemical Equations

Step 1: Write down the chemical equation.
Step 2: List down the atoms (or polyatomic ions) involved in both sides.
Step 3: Count the number of atoms on both sides.
Step 4: Compare both sides and change the coefficients (not subscripts) so that the atoms on the left side are equal to the atoms on the right side.
(Tip: Balance the **Metals** first, then the **Non-Metals**, and then the **Oxygen** atoms and **Hydrogen** atoms.)
Step 5: Double check both sides to make sure the atoms on both sides are equal.

Soluble Salts

Soluble	Insoluble
All nitrates	None
Most sulfates	Lead sulfate, barium sulfate and calcium sulfate
Most chlorides, bromides and iodides	Silver chloride, silver bromide, silver iodide, lead chloride, lead bromide, lead iodide
Sodium carbonate, potassium carbonate, ammonium carbonate	Most other carbonates
Sodium hydroxide, potassium hydroxide, ammonium hydroxide	Most other hydroxides

Uses of Acids

Citric Acid	Used as a sour flavouring agent in food
Hydrochloric Acid	Used as a rust remover
Sulfuric Acid	Used in car batteries
Nitric Acid	Used in fertilisers

Uses of Acids (cont)

Ethanoic Acid	Used as a food preservative
Carbonic Acid	Used in making soft drinks

Uses of Alkalis

Sodium Hydroxide	Used in making soap
Calcium Hydroxide	Used in making toothpaste and to reduce acidity in soil
Aqueous Ammonia	Used in making fertilisers and as a bleaching agent
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Potassium Hydroxide	Used in electroplating and in making cement and plaster
Magnesium Hydroxide	Used as a detergent

Strength of Acids

Strong Acids	Weak Acids
Hydrochloric Acid	Citric Acid
Sulfuric Acid	Tartaric Acid
Nitric Acid	Ethanoic Acid

Strong Acids:
React very fast & vigorously
Ionise completely to produce large amounts of H^+ ions

Weak Acids:
React slowly & less vigorously
Ionise partially to produce small amounts of H^+ ions

Do not confuse the strength of an acid with the concentration of an acid. The strength tells you how many H^+ ions are produced while the concentration tells you how much of an acid is dissolved in water.

Strength of Alkalis

Strong Alkalis	Weak Alkalis
Sodium Hydroxide	Aqueous Ammonia
Potassium Hydroxide	
Calcium Hydroxide	

Strong Alkalis ionise completely to produce large amounts of OH^- ions.
Weak Alkalis ionise partially to produce small amounts of OH^- ions.

How to Carry Out Titration

1. For solid samples, weigh the solid and dissolve in a known volume of solution (usually $100cm^3$).
2. Use a pipette to measure a known volume of the solution (e.g $10cm^3$) and empty into an Erlenmeyer flask.
3. Add a few drops of indicator into the solution.
4. Put the second chemical into a burette. This other solution will react with the synthesised chemical sample in the flask. Often the solution in the burette is an acid or alkali, and it must be of a precise, known concentration.
5. Drop by drop, mix the chemical in the burette into the Erlenmeyer flask until the end point is reached. A colour change indicates the correct amount has been added to react completely with the chemical in the sample.
6. Take note of the volume of the solution added from the burette.