General Chemistry Laboratory Manual

Auburn Riverside HS
Revised Fall 2012

PERIODIC TABLE OF THE ELEMENTS

Table of Selected Radioactive Isotopes

0n	1	(15.3 min) β^-	30Zn	65	(61.88 h) β^+	51Sb	124	(60.20 d) β^+	74W	181	(140 d) EC	91Pa	231	(3.28 × 10^4 y) α
1H	3	(12.26 y) β^-	4Be	7	(53.3 d) EC		125	(2.7 y) β^+		185	(75.1 d) β^+	92U	233	(1.59 × 10^5 y) α
	10	(1.6 × 10^6 y) β^-		72	(14.10 h) β^+	52Te	121 m	(154 d) IT		188	(69 d) β^+		234	(2.44 × 10^5 y) α
6C	11	(20.40 min) β^+		73	(80.3 d) EC		123 m	(119.7 d) IT	75Re	187	(5 × 10^{10} y) β^-		235	(7.04 × 10^8 y) α
	14	(5730 y) β^-	32Ge	68	(275 d) EC		127 m	(109 d) IT	76Os	194	(6.0 y) β^-		236	(2.34 × 10^7 y) α
9F	18	(109.8 min) β^+	33As	73	(80.3 d) EC	53I	129	(1.6 × 10^7 y) β^-	77Ir	192	(74.2 d) β^-, β^+, EC		238	(4.47 × 10^9 y) α
11Na	22	(2.602 y) β^+, EC		74	(17.9 d) β^+, β^+, EC		131	(8.040 d) β^-	78Au	195	(183 d) EC	93Np	236	(1.1 × 10^5 y) EC, β^-
	24	(15.02 h) β^-	34Se	75	(118.5 d) β^+	54Xe	133	(5.25 d) β^-		196	(6.18 d) β^+, EC, β		237	(2.14 × 10^6 y) α
	28	(20.9 h) β^-		79	(6.5 × 10^4 y) β		135	(9.10 h) β^-		198	(2.696 d) β^-		239	(2.356 d) β^-
12Mg	28	(20.9 h) β^-	35Br	82	(35.34 h) β^-	55Cs	134	(2.06 y) β^-		199	(3.15 d) β^-	94Pu	238	(87.75 y) α
13Al	26	(7.2 × 10^5 y) β^+, EC	36Kr	81	(2.1 × 10^5 y) EC		135	(2.9 × 10^6 y) β-	80Hg	203	(46.8 d) β^-		239	(2.41 × 10^4 y) α
14P	32	(14.26 d) β^-		85	(10.72 y) β^-		137	(30.17 y) β^-	81Tl	204	(3.77 y) β^-		240	(6.54 × 10^3 y) α
16S	35	(87.7 d) β^-	37Rb	86	(18.7 d) β^-	56Ba	140	(12.8 d) β^-	82Pb	202	(3 × 10^5 y) EC		242	(3.8 × 10^5 y) α
17Cl	36	(3.01 × 10^5 y) β^-		87	(4.8 × 10^{10} y) β^-	57La	137	(6 × 10^4 y) EC		205	(3 × 10^7 y) EC		244	(8.3 × 10^7 y) α
	38	(37.2 min) β^-	38Sr	90	(28.8 y) β^-		140	(40.3 h) β^-		210	(22.3 y) β^-, α	95Am	241	(432 y) α
18Ar	37	(35.02 d) EC	39Y	88	(106.6 d) β^+, EC	58Ce	144	(284 d) β^-	83Bi	207	(38 y) EC		243	(7.37 × 10^3 y) α
	39	(265 y) β^-	40Zr	93	(1.5 × 10^6 y) β^+	59Pr	142	(19.1 h) β^-		208	(3.7 × 10^5 y) EC	96Cm	242	(163.2 d) α
19K	40	(1.28 × 10^9 y) EC		95	(64.0 d) β^-	60Nd	147	(11.1 d) β^-		210	(5.01 d) β^-, α		244	(18.12 y) α,
	42	(12.36 h) β^-	41Nb	94	(2.0 × 10^4 y) β^-	61Pm	145	(18 y) EC,		210 m	(3 × 10^6 y) α		247	(1.55 × 10^7 y) α
20Ca	45	(165 d) β^-		95	(35.15 d) β^-		147	(2.62 y) β^-	84Po	208	(2.90 y) α		248	(3.5 × 10^5 y) α, SF
21Sc	46	(83.80 d) β^-	42Mo	99	(66.02 h) β^-	62Sm	146	(7 × 10^7 y) α		209	(102 y) α	97Bk	247	(1.4 × 10^7 y) α
24Cr	51	(27.70 d) EC	43Tc	97	(2.6 × 10^6 y) EC		151	(93 y) β^-		210	(138.38 d) α	98Cf	249	(351) α
25Mn	53	(2 × 10^6 y) EC		98	(4.2 × 10^6 y) β^-	63Eu	152	(13 y) β^-, EC, β^-	85At	209	(5.4 h) EC, α		251	(900 y) α
	54	(313.0 d) EC		99	(2.13 × 10^5 y) β^-		154	(8.5 y) β^-		210	(8.1 h) EC	99Es	252	(472 d) α
	56	(2.578 h) β^-	44Ru	106	(367 d) β^-	64Gd	150	(2.1 × 10^6 y) α		211	(7.21 h) EC, α		253	(20.47 d) α
26Fe	59	(44.6 d) β^-	45Rh	101	(3.3 y) EC	65Tb	158	(1.2 × 10^{12} y) EC, β^+	86Rn	222	(3.824 d) α		254	(276 d) α
27Co	56	(78.8d) β^+, EC	46Pd	103	(17.0 d) EC		160	(72.3 d) β^-	87Fr	212	(19.3 min) EC, α	100Fm	255	(20.1 h) α
	57	(270 d) EC		107	(7 × 10^6 y) β^-	66Ho	166 m	(1.2 × 10^3 y) β^-		222	(1.5 min) β^-		257	(100.5 d) α
	58	(71.3 d) β^+, EC	47Ag	108	(127 y) EC	67Ho	170	(1.28 d) β^+		223	(21.8 min) β^-	101Md	258	(55 d) α
	60	(5.272 y) β		110	(252 d) β		171	(1.92 y) β^-	88Ra	226	(1.60 × 10^3 y) α	102No	259	(58 min) α
28Ni	57	(36.0 h) β^+, EC		111	(7.45 d) β^-	68Er	169	(32.0 d) EC	89Ac	227	(21.77 y) β^-	103Lr	260	(3.0 min) α
	59	(8 × 10^4 y) EC	48Cd	109	(453 d) EC		175	(4.19 d) β^-	90Th	228	(1.913 y) α	104Rf	261	(65 s) α
	63	(92 y) β^-	49In	114	(49.51 d) IT	69Tm	167	(9.24 d) EC		230	(7.7 × 10^4 y) α	105Db	262	(40 s) α
29Cu	64	(12.70 h) β^+, β^-, EC	50Sn	121	(76 y) β	70Yb	169	(32.0 d) EC		232	(1.40 × 10^{10} y) α	106Sg	263	(0.9 d) α
							175	(4.19 d) β^-						
						71Lu	176	(3.7 × 10^{10} y) β^-						
						72Ta	182	(115.0 d) β^-						

Selected Radioactive Isotopes

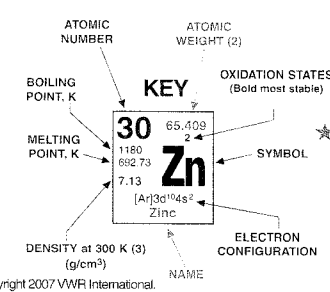
Naturally occurring radioactive isotopes are designated by a mass number in blue (although some are also manufactured). Letter m indicates an isomer of another isotope of the same mass number. Half-lives follow in parentheses, where s, min, h, d, and y stand respectively for seconds, minutes, hours, days, and years. The table includes mainly the longer-lived radioactive isotopes; many others have been prepared. Isotopes known to be radioactive but with half-lives exceeding 10^{12} y have not been included. Symbols describing the principal mode (or modes) of decay are as follows (these processes are generally accompanied by gamma radiation):

α	alpha particle emission
β^-	beta particle (electron) emission
β^+	positron emission
EC	orbital electron capture
IT	isomeric transition from upper to lower isomeric state
SF	spontaneous fission

GROUP 1/IA	2/IIA	3/IIIA	4/IVA	5/VA	6/VIA	7/VIIA	8	9 VIIIA	10	11/IB	12/IIB	13/IIIB	14/IVB	15/VB	16/VIB	17/VIIB	18/VIII
1 1.00794 1,-1 20.28 13.81 0.0899† **H** 1s¹ Hydrogen																	2 4.00260 4.216 0.95 at 26 atm 0.1785† **He** 1s² Helium
3 (6.941) 1 1615 453.7 0.534 **Li** [He]2s¹ Lithium	4 9.012182 2 2744 1560 1.85 **Be** [He]2s² Beryllium											5 10.811 3 4275 2348 2.37 **B** [He]2s²2p¹ Boron	6 12.0107 ±4,2 4675* 3915* 2.26 **C** [He] 2s²2p² Carbon	7 14.0067 ±3,5,4,2 77.344 63.15 1.251† **N** [He] 2s²2p³ Nitrogen	8 15.9994 -2,-1 90.2 54.4 1.429† **O** [He]2s²2p⁴ Oxygen	9 18.99840 -1 85.0 53.55 1.696† **F** [He] 2s²2p⁵ Fluorine	10 20.1797 27.07 24.56 0.900† **Ne** [He] 2s²2p⁶ Neon
11 22.989770 1 1156.1 371.0 0.971 **Na** [Ne]3s¹ Sodium	12 24.3050 2 1363 923 1.74 **Mg** [Ne]3s² Magnesium											13 26.981538 3 2792 933.5 2.6989 **Al** [Ne]3s²3p¹ Aluminum	14 28.0855 4,2 3538 1687 2.33 **Si** [Ne]3s²3p² Silicon	15 30.97376 ±3,5,4 553 317.3 1.82 **P** [Ne]3s²3p³ Phosphorus	16 32.065 6,±2,4, 717.75 388.36 2.07 **S** [Ne]3s²3p⁴ Sulfur	17 35.453 ±1,7,5,3 239.11 171.65 3.214† **Cl** [Ne]3s²3p⁵ Chlorine	18 39.948 87.8 83.8 1.784† **Ar** [Ne]3s²3p⁶ Argon
19 39.0983 1 1033 336.8 0.862 **K** [Ar]4s¹ Potassium	20 40.078 2 1757 1115 1.55 **Ca** [Ar]4s² Calcium	21 44.95591 3 3109 1814 2.99 **Sc** [Ar]3d¹4s² Scandium	22 47.867 4,3,2 3560 1941 4.54 **Ti** [Ar]3d²4s² Titanium	23 50.9415 5,4,3,2 3680 2183 6.11 **V** [Ar]3d³4s² Vanadium	24 51.996 3,6,2 2945 2180 7.19 **Cr** [Ar]3d⁵4s¹ Chromium	25 54.9380 2,3,4,6,7 2234 1519 7.44 **Mn** [Ar]3d⁵4s² Manganese	26 55.845 3,2,6 3134 1811 7.874 **Fe** [Ar]3d⁶4s² Iron	27 58.9332 2,3 3200 1768 8.90 **Co** [Ar]3d⁷4s² Cobalt	28 58.6934 2,3 3186 1728 8.90 **Ni** [Ar]3d⁸4s² Nickel	29 63.546 2,1 2835 1357.8 8.96 **Cu** [Ar]3d¹⁰4s¹ Copper	30 65.409 2 1180 692.68 7.13 **Zn** [Ar]3d¹⁰4s² Zinc	31 69.723 3 2477 302.91 6.095 **Ga** [Ar]3d¹⁰4s²4p¹ Gallium	32 72.64 4 3106 1211.4 5.32 **Ge** [Ar]3d¹⁰4s²4p² Germanium	33 74.9216 ±3,5 876 (subl) 1090 (29 atm) 5.73 **As** [Ar]3d¹⁰4s²4p³ Arsenic	34 78.96 4,6,-2 958 494 4.79 **Se** [Ar]3d¹⁰4s²4p⁴ Selenium	35 79.904 ±1,7,5,3 331.95 265.95 3.12 **Br** [Ar]3d¹⁰4s²4p⁵ Bromine	36 83.80 0,2 119.93 115.8 3.73† **Kr** [Ar]3d¹⁰4s²4p⁶ Krypton
37 85.4678 1 961 312.46 1.532 **Rb** [Kr]5s¹ Rubidium	38 87.62 2 1655 1050 2.54 **Sr** [Kr]5s² Strontium	39 88.9059 3 3618 1795 4.47 **Y** [Kr]4d¹5s² Yttrium	40 91.224 4 4682 2128 6.51 **Zr** [Kr]4d²5s² Zirconium	41 92.90638 5,3,4 5017 2730 8.57 **Nb** [Kr]4d⁴5s¹ Niobium	42 95.94 6,5,4,3,2 4912 2896 10.22 **Mo** [Kr]4d⁵5s¹ Molybdenum	43 (98) 7,5,4 4538 2430 11.5* **Tc** [Kr]4d⁵5s² Technetium	44 101.07 2,3,4,6,8 4423 2607 12.37 **Ru** [Kr]4d⁷5s¹ Ruthenium	45 102.90550 2,3,4 3968 2237 12.41 **Rh** [Kr]4d⁸5s¹ Rhodium	46 106.42 2,4 3236 1828 12.0 **Pd** [Kr]4d¹⁰ Palladium	47 107.8682 1 2435 1234.93 10.50 **Ag** [Kr]4d¹⁰5s¹ Silver	48 112.41 2 1040 594.22 8.65 **Cd** [Kr]4d¹⁰5s² Cadmium	49 114.82 3 2345 429.75 7.31 **In** [Kr]4d¹⁰5s²5p¹ Indium	50 118.710 4,2 2875 505.08 7.31 **Sn** [Kr]4d¹⁰5s²5p² Tin	51 121.760 ±3,5 1860 903.78 6.69 **Sb** [Kr]4d¹⁰5s²5p³ Antimony	52 127.60 4,6,-2 1261 722.66 6.24 **Te** [Kr]4d¹⁰5s²5p⁴ Tellurium	53 126.90447 ±1,5,7 457.51 386.85 4.93 **I** [Kr]4d¹⁰5s²5p⁵ Iodine	54 131.29 0,2,4,6,8 165.11 161.4 5.90† **Xe** [Kr]4d¹⁰5s²5p⁶ Xenon
55 132.90545 1 944 301.54 1.87 **Cs** [Xe]6s¹ Cesium	56 137.327 2 2170 1000 3.5 **Ba** [Xe]6s² Barium	57 138.9055 3 ★ 3737 1191 6.15 **La** [Xe]5d¹6s² Lanthanum	72 178.49 4 4876 2506 13.31 **Hf** [Xe]4f¹⁴5d²6s² Hafnium	73 180.9479 5,4 5730 3290 16.65 **Ta** [Xe]4f¹⁴5d³6s² Tantalum	74 183.84 6,5,4,3,2 5828 3695 19.3 **W** [Xe]4f¹⁴5d⁴6s² Tungsten	75 186.207 7,6,5,4,3,2 5870 3459 21.0 **Re** [Xe]4f¹⁴5d⁵6s² Rhenium	76 190.23 4,8,6,3,2 5285 3306 22.57 **Os** [Xe]4f¹⁴5d⁶6s² Osmium	77 192.217 4,2,6,3 4700 2720 22.42 **Ir** [Xe]4f¹⁴5d⁷6s² Iridium	78 195.08 4,2 4098 2041.55 21.45 **Pt** [Xe]4f¹⁴5d⁹6s¹ Platinum	79 196.96655 3,1 3130 1337.33 19.3 **Au** [Xe]4f¹⁴5d¹⁰6s¹ Gold	80 200.59 2,1 629.88 234.32 13.55 **Hg** [Xe]4f¹⁴5d¹⁰6s² Mercury	81 204.3833 3,1 1746 577 11.85 **Tl** [Xe]4f¹⁴5d¹⁰6s²6p¹ Thallium	82 207.2 2,4 2022 600.61 11.35 **Pb** [Xe]4f¹⁴5d¹⁰6s²6p² Lead	83 208.9804 3,5 1837 544.55 9.75 **Bi** [Xe]4f¹⁴5d¹⁰6s²6p³ Bismuth	84 (209) 4,2,6 — 527 9.3 **Po** [Xe]4f¹⁴5d¹⁰6s²6p⁴ Polonium	85 (210) ±1,7,5,3 610* 575 — **At** [Xe]4f¹⁴5d¹⁰6s²6p⁵ Astatine	86 (222) 0,2 211.4 202 9.73† **Rn** [Xe]4f¹⁴5d¹⁰6s²6p⁶ Radon
87 (223) 1 950* 300 — **Fr** [Rn]7s¹ Francium	88 (226) 2 1413 973 5.0 **Ra** [Rn]7s² Radium	89 (227) 3 ★★ 3470* 1324 10.07 **Ac** [Rn]6d¹7s² Actinium	104 (261) — — **Rf** [Rn]5f¹⁴6d²7s² Rutherfordium	105 (262) — — **Db** [Rn]5f¹⁴6d³7s² Dubnium	106 (266) — — **Sg** [Rn]5f¹⁴6d⁴7s² Seaborgium	107 (264) — — **Bh** [Rn]5f¹⁴6d⁵7s² Bohrium	108 (277) — — **Hs** [Rn]5f¹⁴6d⁶7s² Hassium	109 (268) — — **Mt** [Rn]5f¹⁴6d⁷7s² Meitnerium	110 (269) — — **Ds** [Rn]5f¹⁴6d⁹7s¹ (Darmstadtium)	111 (272) — — **Rg** [Rn]5f¹⁴6d⁹7s² (Roentgenium)	112 (285) — — **Uub** [Rn]5f¹⁴6d¹⁰7s² (Ununbium)	113 (284) — — **Uut** [Rn]5f¹⁴6d¹⁰7s²7p¹ (Ununtrium)	114 (289) — — **Uuq** [Rn]5f¹⁴6d¹⁰7s²7p² (Ununquadium)	115 (288) — — **Uup** [Rn]5f¹⁴6d¹⁰7s²7p³ (Ununpentium)	116 — — **Uuh** [Rn]5f¹⁴6d¹⁰7s²7p⁴ (Ununhexium)	117 (Ununseptium)	118 (Ununoctium)

* Estimated Values

★	58 140.116 3,4 3716 1071 6.77 **Ce** [Xe]4f¹5d¹6s² Cerium	59 140.90765 3,4 3785 1204 6.77 **Pr** [Xe]4f³6s² Praseodymium	60 144.24 3 3347 1294 7.01 **Nd** [Xe]4f⁴6s² Neodymium	61 (145) 3 3273 1315 7.26 **Pm** [Xe]4f⁵6s² Promethium	62 150.36 3,2 2067 1347 7.52 **Sm** [Xe]4f⁶6s² Samarium	63 151.964 3,2 1869 1095 5.24 **Eu** [Xe]4f⁷6s² Europium	64 157.25 3 3546 1586 7.90 **Gd** [Xe]4f⁷5d¹6s² Gadolinium	65 158.92534 3,4 3503 1629 8.23 **Tb** [Xe]4f⁹6s² Terbium	66 162.50 3 2840 1685 8.55 **Dy** [Xe]4f¹⁰6s² Dysprosium	67 164.9303 3 2973 1747 8.795 **Ho** [Xe]4f¹¹6s² Holmium	68 167.26 3 3140 1802 9.07 **Er** [Xe]4f¹²6s² Erbium	69 168.9342 3,2 2223 1818 9.32 **Tm** [Xe]4f¹³6s² Thulium	70 173.04 3,2 1469 1092 6.903 **Yb** [Xe]4f¹⁴6s² Ytterbium	71 174.967 3 3675 1936 9.841 **Lu** [Xe]4f¹⁴5d¹6s² Lutetium
★★	90 232.0381 4 5061 2023 11.72 **Th** [Rn]6d²7s² Thorium	91 231.03588 5,4 4300* 1845 15.4* **Pa** [Rn]5f²6d¹7s² Protactinium	92 238.0289 6,5,4,3,2 4404 1408 18.95 **U** [Rn]5f³6d¹7s² Uranium	93 (237) 5,6,4,3 4175* 917 20.2 **Np** [Rn]5f⁴6d¹7s² Neptunium	94 (244) 4,6,5,3 3505 913 19.84 **Pu** [Rn]5f⁶7s² Plutonium	95 (243) 3,6,5,4,2 2284 1449 13.7 **Am** [Rn]5f⁷7s² Americium	96 (247) 3 — 1620 13.5* **Cm** [Rn]5f⁷6d¹7s² Curium	97 (247) 3,4 — — 14* **Bk** [Rn]5f⁹7s² Berkelium	98 (251) 3,4 — 1170* — **Cf** [Rn]5f¹⁰7s² Californium	99 (252) 3 — 1130* — **Es** [Rn]5f¹¹7s² Einsteinium	100 (257) 3 — 1800* — **Fm** [Rn]5f¹²7s² Fermium	101 (258) 3 — 1100* — **Md** [Rn]5f¹³7s² Mendelevium	102 (259) 2,3 — 1100* — **No** [Rn]5f¹⁴7s² Nobelium	103 (262) 3 — 1900* — **Lr** [Rn]5f¹⁴6d¹7s² Lawrencium



NOTES:

(1) Black — solid. Red — gas. Blue — liquid. Outline — synthetically prepared.

(2) Based upon carbon-12. ( ) indicates most stable or best known isotope.

(3) Entries marked with daggers refer to the gaseous state at 273 K and 1 atm and are given in units of g/l.

The A & B subgroup designations, are those recommended by the International Union of Pure and Applied Chemistry.

Catalog Number WLS-18806

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Side 1

Positive Ions – Cations

1+		2+		3+		4+	
A. Ammonium	NH_4^+	Barium	Ba^{2+}	Aluminum	Al^{3+}	Carbon	C^{4+}
B. Cesium	Cs^+	Beryllium	Be^{2+}	Antimony (III)	Sb^{3+}	Lead (IV)	Pb^{4+}
C. Copper (I)	Cu^+	Cadmium (II)	Cd^{2+}	Bismuth (III)	Bi^{3+}	Silicon	Si^{4+}
D. Gold (I)	Au^+	Calcium	Ca^{2+}	Chromium (III)	Cr^{3+}	Tin (IV)	Sn^{4+}
E. Hydrogen	H^+	Chromium (II)	Cr^{2+}	Cobalt (III)	Co^{3+}		
F. Lithium	Li^+	Cobalt (II)	Co^{2+}	Gallium	Ga^{3+}	5+	
G. Potassium	K^+	Copper (II)	Cu^{2+}	Gold (III)	Au^{3+}	Antimony (V)	Sb^{5+}
H. Rubidium	Rb^+	Iron (II)	Fe^{2+}	Iron (III)	Fe^{3+}	Bismuth (V)	Bi^{5+}
I. Silver	Ag^+	Lead (II)	Pb^{2+}	Manganese (III)	Mn^{3+}		
J. Sodium	Na^+	Magnesium	Mg^{2+}	Nickel (III)	Ni^{3+}		
K.		Manganese (II)	Mn^{2+}				
L.		Mercury (I)	Hg_2^{2+}				
M.		Mercury (II)	Hg^{2+}				
N.		Nickel (II)	Ni^{2+}				
O.		Strontium	Sr^{2+}				
P.		Tin (II)	Sn^{2+}				
Q.		Zinc	Zn^{2+}				

Negative Ions – Anions

-1		-2		-3		-4	
A. Acetate	CH_3COO^-	Carbonate	CO_3^{2-}	A. Arsenide	As^{3-}	A. Carbide	C^{4-}
B. Bromide	Br^-	Chromate	CrO_4^{2-}	B. Nitride	N^{3-}		
C. Chlorate	ClO_3^-	Dichromate	$Cr_2O_7^{2-}$	C. Phosphate	PO_4^{3-}		
D. Chloride	Cl^-	Monohydrogen Phosphate	HPO_4^{2-}	Phosphide	P^{3-}		
E. Chlorite	ClO_2^-	Oxalate	$C_2O_4^{2-}$	Phosphite	PO_3^{3-}		
F. Cyanide	CN^-	Oxide	O^{2-}				
G. Dihydrogen phosphate	$H_2PO_4^-$	Peroxide	O_2^{2-}				
H. Fluoride	F^-	Selenide	Se^{2-}				
I. Hydride	H^-	Silicate	SiO_3^{2-}				
J. Hydrogen carbonate	HCO_3^-	Sulfate	SO_4^{2-}				
K. Hydrogen sulfate	HSO_4^-	Sulfide	S^{2-}				
L. Hydrogen sulfide	HS^-	Sulfite	SO_3^{2-}				
M. Hydrogen sulfite	HSO_3^-	Telluride	Te^{2-}				
N. Hydroxide	OH^-	Thiosulfate	$S_2O_3^{2-}$				
O. Hypochlorite	OCl^-						
P. Iodate	IO_3^-						
Q. Iodide	I^-						
R. Nitrate	NO_3^-						
S. Nitrite	NO_2^-						
T. Perchlorate	ClO_4^-						
U. Permanganate	MnO_4^-						
V. Thiocyanate	SCN^-						

Conclusion

This is the most important part of the lab report.

A Basic Conclusion includes: TS, at least 2 related CDs, ES, CSC

TS **Topic Sentence** is a ***conclusive statement*** which
♦Addresses the Prediction **or** Hypothesis **or**
♦Answers the Investigative Question **or**
♦Outlines the purpose of the lab
TS - In this lab…

CD **Concrete Detail** sentences include concrete DATA from the lab such as
♦High data average of results **and** Low data average of results (**Quantitative**)
♦Statement of what was observed in lab (**Qualitative**)
- *What were the results?*
- *What did you see?*
- *Descriptions of actual observations*

CD - The data show that…

ES – ***Explanatory statement*** -
♦Connects the data results to the prediction / hypothesis &/or investigative question OR
♦Includes derived data (*calculated*) OR
♦Notes possible experimental errors & how these errors may impact the results
- *Did you accomplish your objective?*
- *Did you prove or disprove your hypothesis?*
- *Did the data support your hypothesis?*

ES - From the results…

CSC **Concluding Scientific Concept** –
Is a statement that connects the lab with scientific concepts that have been covered in class.
CSC - This investigation studied… OR

The standard that was covered in this investigation was…

Chunk Writing Conclusion Paragraph

Topic Sentence (TS): **In this lab**, _____

_____

Concrete Detail (CD): **The data shows that…** _____

_____

Concrete Detail (CD): **The data shows that…** _____

_____

Explanatory Statement (ES): **From the results…** _____

_____

Concluding Scientific Concept (CSC): **This investigation studied…**

_____

_____

or

The standard that was covered in this investigation was…_____

_____

Comments / Journal

Proposal of a new hypothesis or explain how the lab can be applied to life. May also include personal comments or question to teacher regarding the lab and its findings.

This may show that…_____**because…** _____

_____

Or

This can be applied to my life… _____

_____

Table of Contents

Equipment Frenzy

Working together, match the equipment at your lab station with the cards provided. Then, complete the worksheet with the following steps:

1. Sketch the item(s) next to its name.
2. Provide your best guess(es) as to what the item(s) are used for

Name	Sketch	Uses
Apron		
Beaker		
Beaker Tongs		
Book		
Bottle		
Buret		
Buret Clamp		
Ceramic Triangle		

Name	Sketch	Uses
Disposable Pipet		
Dropper Bottle		
Electronic Balance		
Erlenmeyer Flask		
Evaporating Dish		
Forceps		
Glass Stirring Rod		
Glassware Brushes		
Graduated Cylinder		

Name	Sketch	Uses
Mortar & Pestle		
Plastic Tray		
Reflux Condenser		
Ring Stand		
Safety Goggles		
Scoopula		
Spatula		
Test Tube		
Test Tube Clamp		
Test Tube Holder		

Name	Sketch	Uses
Thermometer		
Watch Glass		
Weigh Boat		
Wire Gauze		

Uncertainty and Measurement

Introduction

Suppose you have a summer job monitoring the pollution in a local lake. You are interested to collect three 100 ml water samples at certain locations at set times each day. To each sample, you add 5 ml of a coloring agent that reacts and changes color intensity in proportion to the amount of pollutant in the water. You then check each sample with an instrument that detects color intensity and gives a ***quantitative***, or numerical, measure of the amount of pollutant in each sample. Unfortunately, your measurements of similar samples vary by 10 to 20 percent. How could you increase the accuracy and precision of your measurements?

Every measurement has a n ***uncertainty***, or built-in error. This error is due to the limitations in the measurement scale, the manufacturing process, and the ability of the human eye to detect small differences. For example, when measuring volume with a graduated cylinder, the width of scale lines, variations in glass thickness, and the slight changes in the angle of sight when reading the scale, are some of the factors that cause uncertainty. Because of this uncertainty, no measurement should be thought of as an exact value, but rather as a value within a range that varies with the uncertainty. For example, the uncertainty of a volume measurement made with a 100 ml graduated cylinder may be ±0.5 ml. Thus, if you measured 100 ml of water, the actual volume would be 100.0 ±0.5 ml or within a range of 99.5 ml to 100.5 ml. Although ±0.5 ml represents only a 0.5% error for a 100.0 ml measurement, it becomes a much larger error of ±10% when you measure a smaller quantity, such as 5.0 ml.

There are two important lessons you should learn about making measurements. First, you should familiarize yourself with the scale of each piece of lab equipment and learn to read each scale as accurately as possible. Second, you should know the uncertainty of your measurements, because your results cannot me more accurate than the built-in error allows.

In this laboratory investigation, you will become familiar with the measurement scale of graduated cylinders and a thermometer. Then you will determine the uncertainty of measurements made with this equipment. If you really do get a job monitoring water pollution, you will know how to increase the accuracy and precision of your measurements so that they are scientifically useful.

Pre-lab questions

1. Why is it important to wear eye protection at all times in the chemistry lab, even you are not using open flame or dangerous chemicals?
2. In the diagram of the graduated cylinder shown in figure 1-2, what fraction of a ml does each division, or increment, between the 1 ml markings represent?
3. What is the volume of the liquid shown in figure 1-2?
4. Which do you think will have a more predictable impact on the measurements you make in lab, human error or uncertainty in the measurement scales of the lab equipment? Explain.

Materials

Goggles	apron	hot plate
Electronic balance	dropper	plastic tub
Graduated cylinder, 10 ml	Celsius thermometer	ice
Graduated cylinder, 100 ml	Hot hands	

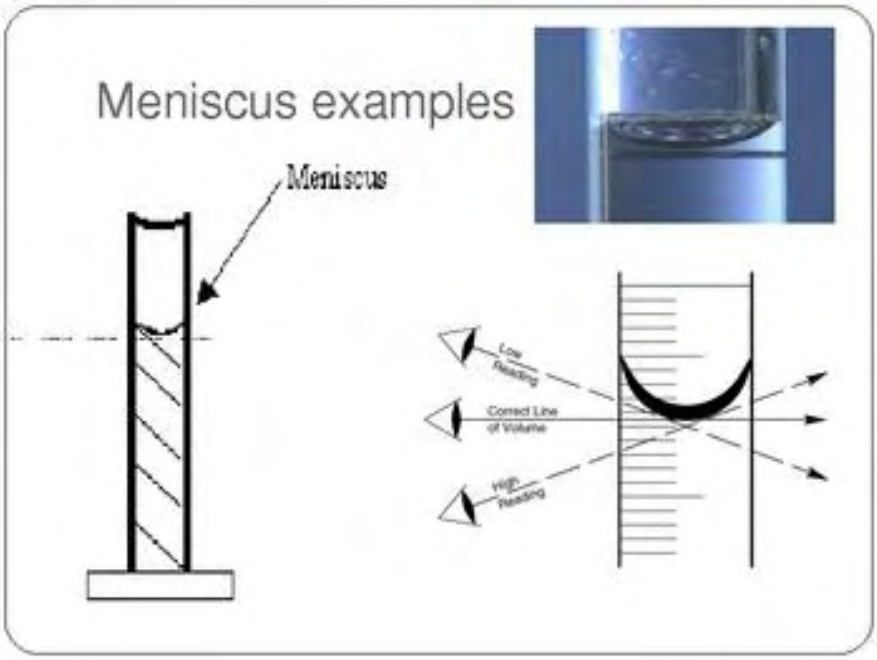


Figure 1-1

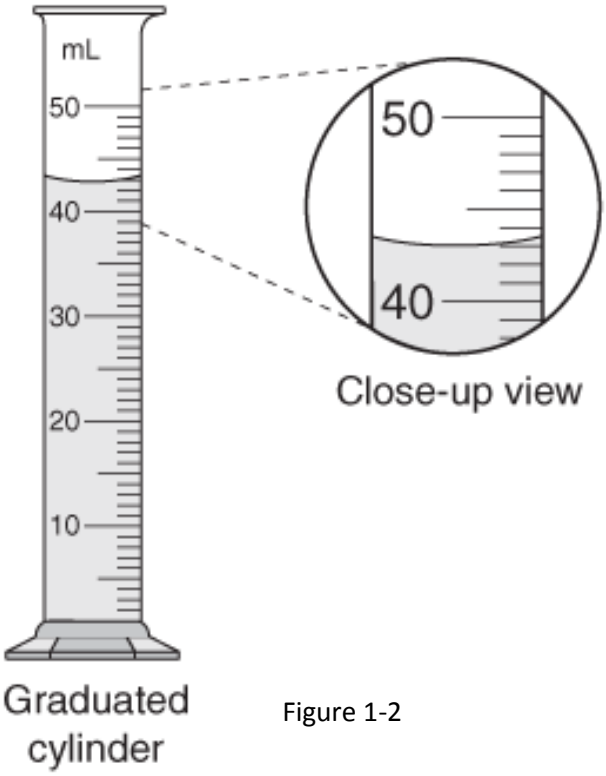


Figure 1-2

Procedure

Uncertainty of graduated cylinders

1. Using an electronic balance, measure and record the mass of a dry 10 ml graduated cylinder and a dry 100 ml graduated cylinder.
2. Record the volume represented by the smallest increment on each of the cylinders. Also, determine and record the volume represented by ½ and ⅕ of the smallest volume increment.
3. Use a dropper to add 10.0 ml distilled water to each cylinder. Add the last few drops to each cylinder carefully so that the bottom curve of the meniscus is on the 10.0 ml mark. Figure 1-1 shows you where to position your sight line so that you obtain an accurate reading.
4. With the electronic balance, measure and record the mass of each cylinder containing 10.0 ml of water.
5. Subtract the mass of the empty cylinder to find the mass of water in the 10 ml cylinder. Pour the water down the sink, dry out the cylinder and repeat 1-5 for the 100 ml cylinder.

Uncertainty of a thermometer

6. Obtain a Celsius thermometer. Determine and record the temperature represented by the smallest scale increment. Also determine and record the temperature represented by ½ and ⅕ of the smallest scale increment.
7. Place the thermometer in the boiling water using a thermometer holder so the tip of thermometer is off of the bottom of the beaker until it shows a constant temperature. Remove the thermometer and quickly record the temperature, estimating to the tenth of a degree.
8. Place you thermometer in a beaker of ice water until it shows a constant temperature. Remove the thermometer and quickly record the temperature, estimating to the tenth of a degree.
9. Clean up all equipment and lab area.

Data table

Graduated cylinders	10 ml	100 ml
Smallest volume scale increment	_____	_____
½ of the smallest volume increment	_____	_____
⅕ of the smallest volume increment	_____	_____
Mass of empty cylinder	_____	_____
Mass of cylinder with 10.0 ml of water	_____	_____
Mass of 10.0 ml of water	_____	_____

Thermometer	
Smallest temperature scale increment	_____
½ of the smallest temperature increment	_____
⅕ of the smallest temperature increment	_____
Temperature of boiling water	_____
Temperature of freezing water	_____

Calculation

Using cylinder class data:

10 ml water, 10 ml graduated cylinder Average: _____

1. 5.
2. 6.
3. 7.
4. 8.

10 ml water, 100 ml graduated cylinder Average: _____

1. 5.
2. 6.
3. 7.
4. 8.

Take the difference (absolute value) between each mass and the average:

10 ml cylinder	100 ml cylinder	
1.	1.	Avg difference 10 ml _____
2.	2.	Rounded 1 significant digit
3.	3.	
4.	4.	Avg difference 100 ml _____
5.	5.	Rounded to 1 significant digit
6.	6.	
7.	7.	
8.	8.	

This is the practical uncertainty since 1.00 g of water at 20°C has a volume of 1.00 ml.

Report the final measurement with the uncertainty included:

_____

_____

Using temperature class data:

Boiling temperature Average: _____

1. 5.
2. 6.
3. 7.
4. 8.

Freezing temperature Average: _____

1. 5.
2. 6.
3. 7.
4. 8.

Take the difference (absolute value) between each mass and the average:

Boiling temperature	Freezing temperature	
1.	1.	Avg difference B.P _____
2.	2.	Rounded 1 significant digit
3.	3.	
4.	4.	Avg difference F.P _____
5.	5.	Rounded to 1 significant digit
6.	6.	
7.	7.	
8.	8.	

Report the final measurement with the uncertainty included:

_____

_____

Analysis

Which did you find to have a smaller uncertainty, 10 ml or 100 ml graduated cylinder? Give a reason why one has a smaller uncertainty.

Do the uncertainties you calculated more closely match the size of ½ or ⅕ the smallest increment?

Suppose a student asks your advice about how to measure 9 ml of a liquid as accurately as possible using a graduated cylinder. Would you recommend a 10 ml or 100 ml graduated cylinder? Support your answer using the results.

Name _____
Date _____
Period_____

Density of Soda

Part A: The Tank

Make an observation about the two sodas in the tank. Come up with three possible reasons to explain what you see.

_____

_____

_____

_____

_____

_____

_____

Part B: The Cans

Determine the density of the full cans.

Coke

Mass:_____

Volume:_____

Density = _____

Diet Coke

Mass:_____

Volume:_____

Density = _____

Part C: The Soda

Mass(Empty):_____

$M_{full} - M_{empty}$ = _____

Volume = _____

Density = _____

Mass(Empty):_____

$M_{full} - M_{empty}$ = _____

Volume = _____

Density = _____

Name _____
Date _____
Period_____

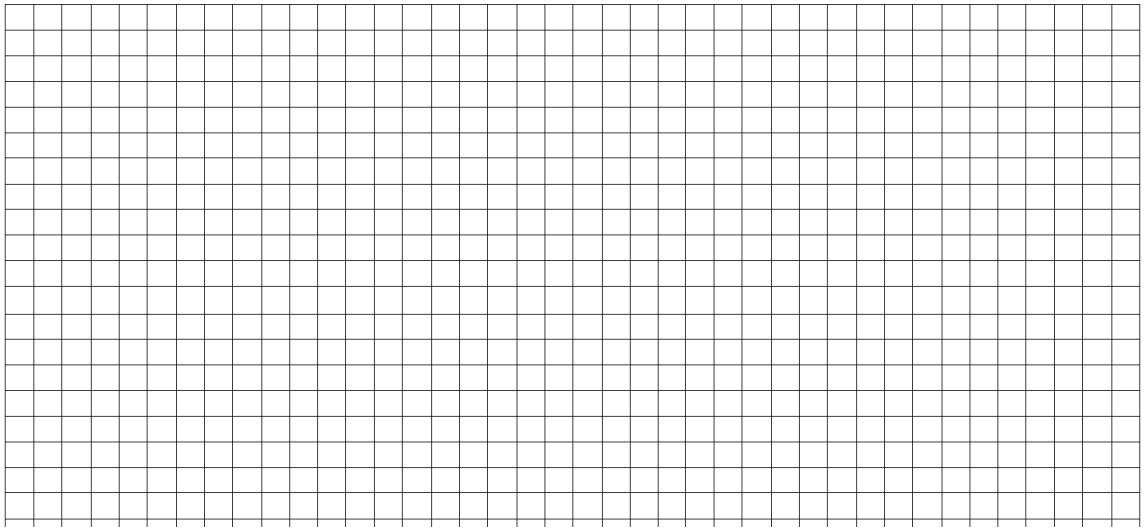
Part D: The Graph

Materials: pipette 10 ml graduated cylinder balance Coke Diet Coke

1. Place a 10 ml graduated cylinder on an electronic balance and zero out.
2. Add approximately 2 ml of coke to the graduated cylinder.
3. Record the mass and the volume (it does not need to be exactly 2 ml, but you do need to read an exact volume)
4. Add approximately 2 ml more.
5. Record mass and volume again.
6. Repeat until you reach 10 ml.
7. Clean rinse and dry the cylinder.
8. Repeat 1-7 using Diet Coke.
9. Graph the data on the table. Y-axis mass, x-axis volume.
10. Determine the slope of the best fit line to the data.

Coke	Mass	Volume
1		
2		
3		
4		
5		

Diet Coke	Mass	Volume
1		
2		
3		
4		
5		



Slope = _____
Questions:

Name _____
Date _____
Period_____

1. The density of pure water is 1.0 g/cm^3. Use this information to explain what you saw in Part A. In your explanation, refer to your calculated density for a full can.

2. Compare the densities of soda you determined: calculation and graph (just the soda). Which method do you think is more accurate? Explain why.

3. What have you determined to be the most significant factor for the difference in densities to explain why one soda floats and why one sinks? Explain.

PERIODIC TABLE OF THE ELEMENTS

This activity is designed to aid you in studying – the tables you create can be used as study guides...

PERIODIC TABLE "A" – Instructions:

1. Subtitle this periodic table "Groups and Blocks of the Periodic Table"
2. Number the groups and periods
3. Make a LEGEND for the following information
    a. Use colors for these
        i. Alkali metals
        ii. Alkaline Earth Metals
        iii. Metalloids
        iv. Nonmetals
        v. Halogens
        vi. Noble Gases
        vii. Transition Metals
        viii. Inner Transition Metals - Lanthanide Series
        ix. Inner Transition Metals - Actinide Series
    b. Use lines for these
        i. *s-block*
        ii. *p-block*
        iii. *d-block*
        iv. *f-block*
4. What is the relationship between the electron configuration of an element and the period in which that element appears in the table?

5. What information is provided by the specific block location of an element?

PERIODIC TABLE "B" – Instructions:

1. Subtitle this periodic table "Metal/Nonmetal/Metalloid, Valence Electrons/Oxidation Number, Ionization Energies, and Electronegativity"
2. Make a LEGEND for the following information (use colors)
    a. Metals
    b. Nonmetals
    c. Metalloids
3. At the top of each group (in the first box) give the most probable valence number. (Do this for groups 1&2 and 13-18 only)
4. What are valence electrons and where are they generally located in the atom?

5. What causes chemical compounds to form?

6. What is meant by atomic radius and what causes atomic radii to decrease?

7. Using lines illustrate the trend of the first ionization energies of main-group elements across a period and down a group. What causes this trend?

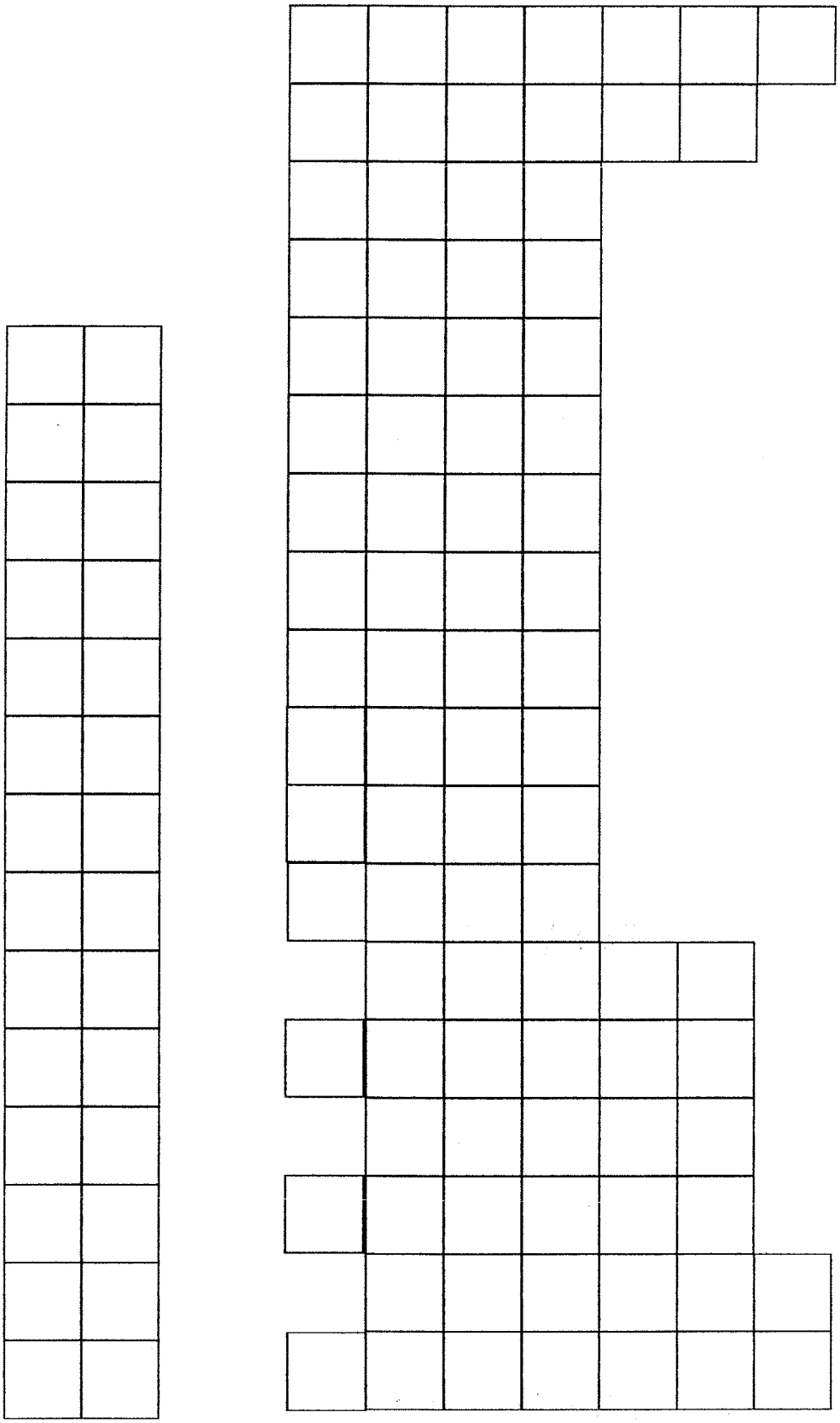
8. Using lines illustrate the trend of electronegativity of main-group elements across a period and down a group. What causes this trend?

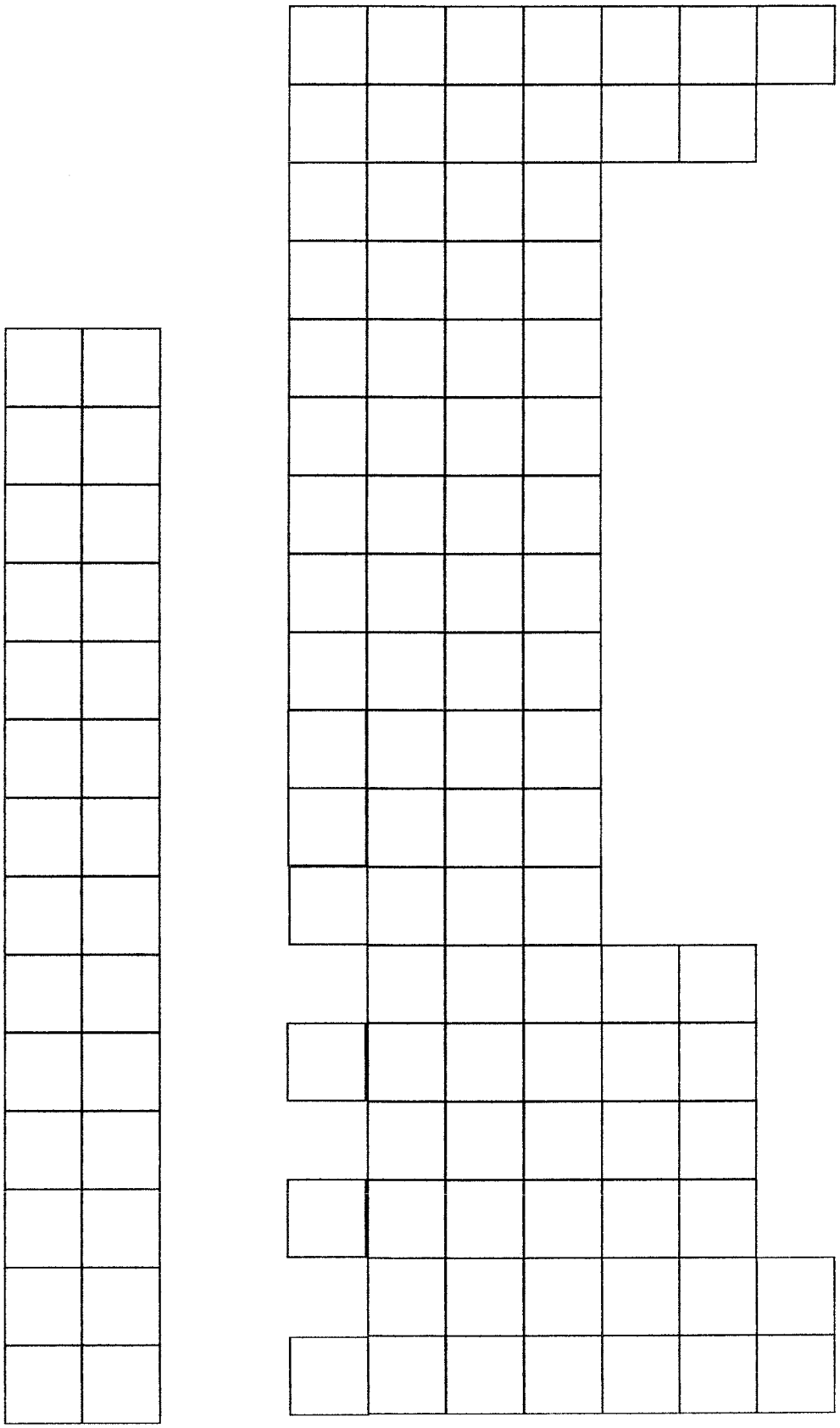
9. Define each term:
    a. Ion

    b. Ionization

    c. Ionization energy

10. What is electronegativity?





Name _____ Date _____ Class _____

Simulating a Cold Pack

Lab 4
APPLICATION

Text reference: ***Chapter 2***

Introduction

Suppose you are on a hike and you sprain your ankle. The immediate application of a cold pack would be a wise first-aid practice. Injuries such as a sprained ankle are accompanied by an increase in blood flow to the affected area, which brings excess heat and contributes to swelling. A cold pack is much colder than your injured ankle, so it removes some of the heat, causes blood vessels to constrict, and reduces swelling, inflammation, and pain.

How exactly does a cold pack work? An instant cold pack, shown in Figure 4–1, usually consists of a tough plastic bag with two compounds inside: water and a salt such as ammonium nitrate (NH_4NO_3), a common lawn fertilizer. The water is sealed inside a fragile inner bag to keep it separated from the ammonium nitrate.

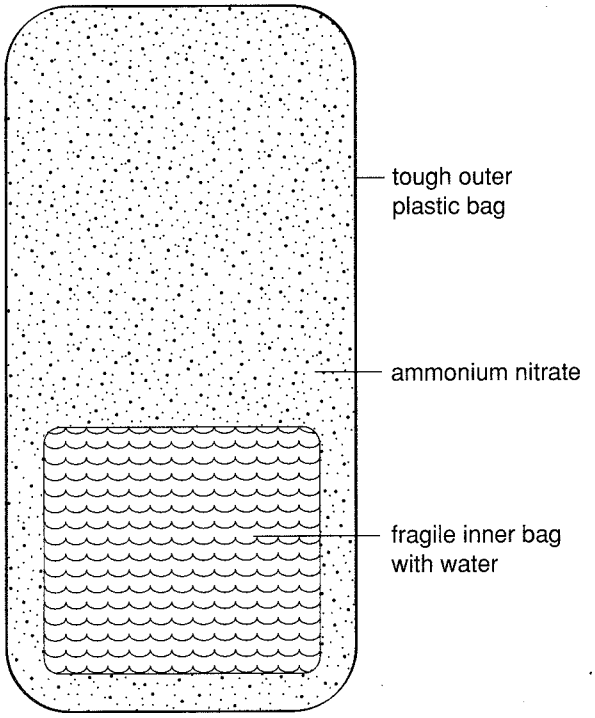


Figure 4–1

When the cold pack is needed, the ammonium nitrate is brought into contact with the water by squeezing the pack until the fragile inner container pops open. As the ammonium nitrate dissolves in the water, a subtle chemical change occurs. The water breaks the solid ammonium nitrate into positively and negatively charged particles (ions).

Lab 4 APPLICATION

Name _____

Chemical changes always involve changes in energy. Often, heat is released, which may be detected as an increase in temperature. Other times, heat energy is absorbed, which results in a decrease in temperature.

In this investigation, you will experiment with the materials that make up an instant cold pack. You will combine water and ammonium nitrate in an insulated cup. The cup will prevent any heat exchange with the environment while the dissolving process removes heat from the water. In order to have a useful cold pack, there must be more than enough solid ammonium nitrate present to reduce the temperature of the liquid in the bag to near zero and keep it cold for an additional 10 to 15 minutes. You will determine the amount of ammonium nitrate solid necessary to lower the temperature to that of melting ice or below and to maintain that temperature for at least 10 minutes.

Pre-Lab Discussion

Read the entire laboratory investigation and the relevant pages of your textbook. Then answer the questions that follow.

1. Define *heat.* _____

_____

2. Define *temperature.* _____

_____

3. In your experience, in what direction does heat exchange occur?

_____

4. Why are plastic foam cups used in this investigation? _____

_____

_____

5. Why should ammonium nitrate not be exposed to an open flame or to temperatures above 250°C? _____

_____

6. Why should clothing splashed with ammonium nitrate solution be rinsed immediately with water? _____

_____

Problem

How can an effective cold pack be made from ammonium nitrate and water?

Materials

chemical splash goggles
laboratory apron
graduated cylinder, 100-mL
tap water
plastic foam cup
4 large pieces weighing paper
laboratory balance
spatula
ammonium nitrate (NH_4NO_3)
thermometer
stirring rod
clock

Name _____

Lab 4
APPLICATION

Safety

Wear your goggles and lab apron at all times during the investigation. Ammonium nitrate is poisonous and can burn or explode when dry or if it is exposed to temperatures above 250°C. Do not expose ammonium nitrate to fire or store in a hot environment. Clothing splashed with ammonium nitrate solution becomes flammable when dried. If ammonium nitrate accidentally comes in contact with your skin or clothing, rinse it off with large quantities of water and inform your teacher immediately. Dispose of the ammonium nitrate solution as instructed by your teacher.

Note the caution alert symbols here and with certain steps of the Procedure. Refer to page *xi* for the specific precautions associated with each symbol.

Procedure

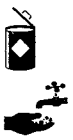
1. Put on your goggles and lab apron. Fill the graduated cylinder with 50 mL of tap water that is at room temperature, and pour the water into the plastic foam cup.

2. Place a weighing paper on the laboratory balance and determine the mass of the paper. Then use a spatula to measure out a mass of 15.0 g of ammonium nitrate, allowing for the mass of the paper. Repeat this step three more times, so that you have four 15.0-g portions of ammonium nitrate. **CAUTION:** *Ammonium nitrate is poisonous and may explode or burn when dry. Avoid contact with skin and clothing. If contact occurs, rinse with large quantities of water.*

3. Using the thermometer, measure the temperature of the water and record this value in the data table.

4. Add one of the 15.0-g portions of ammonium nitrate to the cup and stir it slowly with the stirring rod until it is dissolved. Measure and record the temperature of the water. Also record your observations, such as how much time it takes for the ammonium nitrate to dissolve and how quickly the temperature of the solution changes.

5. Repeat Step 4 with each of the remaining 15.0-g portions of ammonium nitrate until enough of the compound is present to maintain the temperature of the water at 0°C or below for at least 10 minutes.



6. Dispose of the ammonium nitrate solution according to your teacher's instructions. Clean up your work area and wash your hands before leaving the laboratory.

Lab 4 APPLICATION

Name _____

Observations

Data Table

Mass of NH_4NO_3 in Solution (g)	Temp. (°C)	Change in Temp. (°C)	Observations
0.0			
15.0			
30.0			
45.0			
60.0			

Calculations

1. Calculate the change in temperature that occurs after each 15.0-g portion of ammonium nitrate is added to the water, and write these values in the Data Table.

Critical Thinking: Analysis and Conclusions

1. At what point did the largest change in temperature occur? How can this observation be explained? *(Interpreting data)* _____

2. Identify any patterns that you observed as the portions of ammonium nitrate were added to the water. *(Making comparisons)* _____

3. What do you think would happen if you tried to dissolve another 15.0 g of ammonium nitrate in the water? Explain your reasoning. *(Drawing conclusions)* _____

Critical Thinking: Applications

1. In this investigation, the insulated cup prevents the cold solution from being warmed by the outside environment. Would you expect a commercial cold pack to remain as cold for the same length of time if it was applied to an injured ankle? Explain. *(Applying concepts)*

Name _____

Lab 4 APPLICATION

2. Predict what happens to any undissolved ammonium nitrate in the cold pack as heat is absorbed. *(Making predictions)* _____

3. One of the dangers of using a cold pack made of ice—especially one taken from a freezer—is frostbite. In light of the temperatures reached by the addition of the final portions of ammonium nitrate, would you expect the danger of frostbite to exist with a cold pack made from ammonium nitrate? Explain. *(Making judgments)* _____

Going Further

1. Describe how you would design an instant cold pack from common household items, ammonium nitrate, and water. Include a list of the materials needed, directions for using the cold pack, and the appropriate safety warnings. Make a sketch of your design. (Note: You can calculate the ratio of the mass of ammonium nitrate per gram of water used in the investigation, and use the ratio to specify the amount of each compound needed in your design.)
2. Ammonium nitrate is commonly used as a fertilizer. You might think that it could be disposed of simply by adding it to soil or water. Consult an environmentalist or do library research to determine what problems might be created by dumping large amounts of ammonium nitrate into the environment.

1. Recall the different states of matter (Solid, Liquid, Gas). How do the water molecules differ in the liquid and gas states? Explain or draw.

a) Label the phase changes on your heating curve above.

b) What happens to the molecules as they begin to boil?

c) Did you stop adding heat at any point during the lab? As heat was added what happened to the energy of system?

d) If you were always adding heat, then why did the temperature trend change?

2. When you start the experiment, the hot plate is putting heat energy into the system but the curve is almost flat, what is the heat energy doing if it isn't increasing the temperature?

3. What happens to the energy being absorbed from the heat source? Use the heating curve and your knowledge of atoms to explain.

Measuring the Specific Heats of Metals

Introduction

Heat is a form of energy that cannot be measured directly. However, changes in heat can be determined by measuring changes in temperature. When a substance is heated, the heat gained, *q,* depends upon three important factors; the mass, m, of the substance in grams, the specific heat of the substance, cp, and the substance's change in temperature, Δt. As discussed previously, to calculate the amount of heat absorbed or lost by a substance, the following equation is used:

$$q = m \times cp \times \Delta t$$

The specific heat of a substance is an intrinsic physical property, reflecting the type of bonding and intermolecular forces in the substance. In this investigation, you will measure the specific heats of several metals. First, you will need to make and calibrate a calorimeter, an instrument for measuring heat changes. You will start with a predetermined mass of water in your calorimeter. When the heated metal is added to the water, the change in temperature for the water will be measured. Using the specific heat of water (4.184 J/g•C°), you will be able to calculate the amount of heat gained by the water. This value will be equal to the amount of heat lost by the metal. If the temperature change and mass of the metal are known, its specific heat can be determined.

$$\mathbf{q_{metal} = q_{calorimeter}}$$
$$\mathbf{m_{metal} \times \Delta T_{metal} \times Cp_{metal} = m_{water} \times \Delta T_{water} \times Cp_{water}}$$

Objectives

- **Relate** measurements of temperature to changes in heat content
- **Calculate** the specific heats of several metals
- **Determine** the identity of an unknown metal
- **Infer information about** the strength of intermolecular forces

Materials

- 2 plastic foam cups for calorimeter
- thermometer
- 100 mL graduated cylinder
- two 400 mL beakers
- hot plate
- large test tube
- balance
- tongs to handle hot beakers and metal
- unknown metal sample
- boiling chips

Metal Specific Heat Tests

	H_2O	Metal
Initial Temp.		
Final Temp.		
Change in Temp.		
Mass		

1. Put 75.0 mL of DI water into the calorimeter and allow it to come to room temperature. Record the mass and temperature of the water in the data table.

2. Fill a 400 mL beaker with water and add some boiling chips. Use enough metal to fit into the large test tube but remain below the surface of the water when the test tube is placed into the beaker. You need to use as much metal as possible. **Measure the mass of the metal as precisely as possible before placing it in the test tube.** Record the mass in the data table.

3. Clamp the test tube to the ring stand, as shown in Figure 3-B. Set the beaker upon the wire gauze on the ring clamp. Lower the test tube clamp so that the test tube is lowered into the beaker of water. Heat it until the water

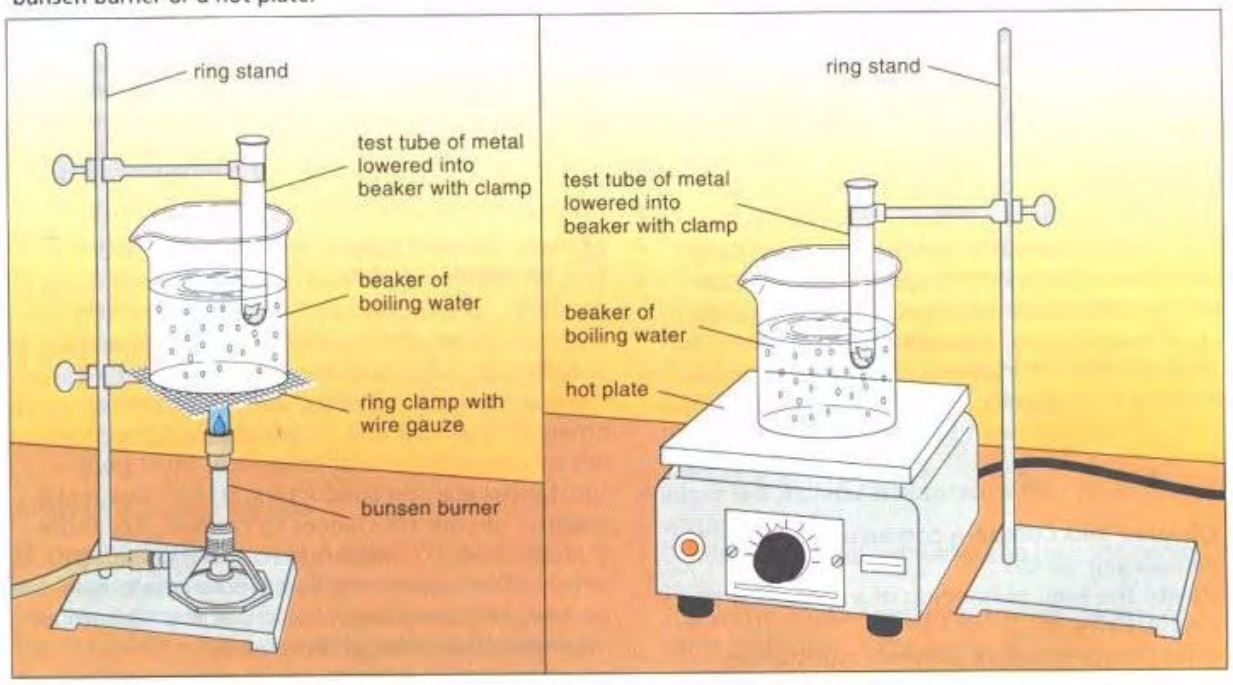
begins to boil. Allow the boiling to continue for 5 minutes. You can now assume the metal is at the boiling temperature of water. Record this temperature in the second data table.

4. Transfer the metal to the calorimeter, holding the test tube with the clamp.
   Be careful not to burn your skin with the hot water or the metal!
   Put the top on the calorimeter. Mix or stir for 30 seconds. Now, record the highest temperature of the water.

5. Repeat Steps 6-10 as needed for other metals, including the unknown metal.

6. Check with your teacher for instructions on how to dispose of your chemicals when finished. Remember to wash your hands thoroughly before leaving the laboratory.

Figure 3-B: The metal samples are each heated by a hot-water bath using either a bunsen burner or a hot-plate.



Analysis and Interpretation

6. **Inferring Relationships:** For the second part of the investigation, write an equation for the relationship between the heat lost by a metal, the heat gained by the water and the heat gained by the calorimeter. Solve for the specific heat of the metal by substituting the equation for specific heat ($q = m \times c_p \times \Delta T$) into the first equation. Show all work for each metal

Conclusions

1. **Inferring Conclusions:** Finish filling in the second set of data tables and use the relationship from Analysis & Interpretation #6 to calculate the specific heat of each of the metals tested.

2. **Evaluating Conclusions:** Compare your value for the unknown metal to the values given in class. What do you think your unknown is made of?

3. **Evaluating Conclusions:** Compare your values for specific heats to the table values that can be found in a handbook such as the CRC *Handbook of* Chemistry *and Physics* or *Lange's Handbook of* Chemistry. (You may need to convert the values given from calories to joules to use the handbook values. The conversion factor is the same as the specific heat of water: 1 cal = 4.184 J.) Calculate your percent error.

4. **Analyzing Conclusions:** How do the specific heats of the metals compare to the specific heat of water? What does that imply about the amount of heat needed to heat a certain mass of metal or water to the same temperature? Which one would absorb heat better?

5. **Organizing Ideas:** Are higher values or lower values better for the specific heat of a calorimeter? Why?

Metal	Specific Heat Capacity - c_p		
	(kJ/kg K)	*(kcal/kg^{o}C)*	*(Btu/lb$_m$ ^{o}F)*
Aluminum	0.91	0.22	0.22
Antimony	0.21	0.05	0.05
Beryllium	1.83	0.436	0.436
Bismuth	0.13	0.03	0.03
Cadmium	0.23	0.055	0.055
Carbon Steel	0.49	0.12	0.12
Cast Iron	0.46	0.11	0.11
Chromium	0.46	0.11	0.11
Cobalt	0.42	0.1	0.1
Copper	0.39	0.092	0.09
Gold	0.13	0.031	0.03
Iridium	0.13	0.031	0.31
Iron	0.46	0.108	0.11
Lead	0.13	0.031	0.03
Magnesium	1.05	0.243	0.25
Manganese	0.48	0.114	0.114
Mercury	0.14	0.033	0.03
Molybdenum	0.25	0.06	0.06
Nickel	0.54	0.106	0.13
Niobium (Columbium)	0.27	0.064	0.064
Osmium	0.13	0.031	0.031
Platinum	0.13	0.032	0.03
Plutonium	0.13	0.032	0.032
Potassium	0.75	0.180	0.180
Rhodium	0.24	0.058	0.058
Selenium	0.32	0.077	0.077
Silicon	0.71	0.17	0.17
Silver	0.23	0.057	0.057
Sodium	1.21	0.29	0.29
Tantalum	0.14	0.034	0.34
Thorium	0.13	0.03	0.03
Tin	0.21	0.054	0.05
Titanium	0.54	0.125	0.13
Tungsten	0.13	0.032	0.03
Uranium	0.12	0.028	0.028
Vanadium	0.39	0.116	0.116
Zinc	0.39	0.093	0.09
Wrought Iron	0.50	0.12	0.12

Name _____ Date _____ Class _____

Heat of Solution

Lab 34

Text reference: ***Chapter 12***

Introduction

Whenever bonds are formed or broken in a chemical reaction, energy is transferred. As a solid dissolves in water, its bonds are broken, and a change in temperature is usually observed. If energy is absorbed from the solvent when a solid dissolves, the system gets colder—the reaction is endothermic, and it has a positive enthalpy change. On the other hand, if energy is released, the system gets warmer—the reaction is exothermic, and it has a negative enthalpy change.

The molar heat of solution of a compound is the heat transferred when one mole of the compound (the solute) dissolves in a solvent. That change cannot be measured in isolation, but it can be determined if the amount of solute dissolved, the amount of solvent used, and the temperature change of the solution are all known.

In this investigation, you will explore the energy changes that take place when various solids (the solutes) dissolve in water (the solvent). You will dissolve known quantities of several solids in different 100-mL samples of water and measure the temperature changes as they dissolve. From these data, the molar heat of solution for each solid will be found.

Pre-Lab Discussion

Read the entire laboratory investigation and the relevant pages of your textbook. Then answer the questions that follow.

1. What is the definition of a chemical bond? _____

   _____

2. What is the difference between heat and temperature? _____

   _____

   _____

3. When the reactants get colder in an endothermic reaction, what has happened to the heat energy? _____

   _____

   _____

4. Is the change in enthalpy positive or negative for an exothermic reaction? Explain. _____

   _____

   _____

Name _____

5. Why is a plastic foam cup used instead of a beaker? For what piece of equipment is this cup a substitute? _____

_____

_____

6. Why should you not stir the solution with the thermometer? What should be used? _____

_____

Problem

How can you measure the molar heat of solution of solids?

Materials

chemical splash goggles
laboratory apron
latex gloves
graduated cylinder, 100-mL
distilled water
large plastic foam cup
thermometer
laboratory balance
sodium thiosulfate pentahydrate ($Na_2S_2O_3 \cdot 5H_2O$)
4 pieces of weighing paper
stirring rod
sodium hydroxide pellets (NaOH)
ammonium chloride (NH_4Cl)
calcium chloride ($CaCl_2$)

Safety

Wear your goggles and lab apron at all times during the investigation. Sodium hydroxide is a very strong base and is very caustic. Wear latex gloves when working with the sodium hydroxide crystals, and do not touch them with your hands. The other solutions may also be irritating to skin. Avoid contact with them. Clean up any spills with plenty of cold water.

Note the caution alert symbols here and with certain steps of the Procedure. Refer to page *xi* for the specific precautions associated with each symbol.

Procedure

1. Put on your goggles and lab apron. Measure 100.0 mL of distilled water at room temperature and pour it into the plastic foam cup. Record the temperature of the water in the Data Table. Do not remove the thermometer from the cup, but be careful that it does not tip the cup over.
2. Using your laboratory balance, measure out 5–10 grams of the sodium thiosulfate on a piece of paper. Record the mass used to 0.01 g. (Make sure you take the mass of the paper into account.)

3. Without removing the thermometer from the cup, shake the sodium thiosulfate from the paper into the water and stir gently with the stirring rod until the solid is completely dissolved. **CAUTION:** *All of the solutions in this investigation are irritating to skin. Avoid contact with them.*
4. Make sure that the bulb of the thermometer is fully immersed in the liquid. If the temperature rises, record the highest temperature reached by the solution. If the temperature falls, record the lowest temperature.
5. Dispose of the solution by pouring it down the drain, followed by plenty of water. Rinse and dry the cup.

6. Put on your latex gloves. Repeat Steps 1–5, using 2–4 grams of sodium hydroxide pellets as your solute. **CAUTION:** *Sodium hydroxide is a very strong base and is very caustic. Wear latex gloves when working with the sodium hydroxide pellets, and do not touch them with your hands. Clean up any spills with plenty of cold water.*
7. Repeat Steps 1–5, using 8–12 grams of ammonium chloride as your solute.
8. Repeat Steps 1–5, using 8–12 grams of calcium chloride as your solute.

9. Clean up your work area and wash your hands before leaving the laboratory.

Observations

DATA TABLE

Solute	Solute Mass (g)	Solution Mass (g)	Initial T (°C)	Final T (°C)	ΔT (°C)
$Na_2S_2O_3 \cdot 5H_2O$					
NaOH					
NH_4Cl					
$CaCl_2$					

Calculations

1. ΔT is the change from initial temperature to final temperature ($\Delta T = T_f - T_i$). Calculate ΔT for each reaction, and enter the values in the Data Table.
2. Find the total mass of solution that changed temperature in each reaction, and write the values in the Data Table. (Assume that 100.0 mL of water has a mass of 100.0 g.)
3. Calculate the energy (in joules) absorbed or released by the solution in each reaction. Use the absolute value for ΔT in the following equation.

$$\text{energy} = (\text{mass}) \times (\Delta T) \times (\text{specific heat of solution})$$

Name _____

The specific heat of a solution is slightly different from the specific heat of pure water. In this investigation, assume it is 3.34 J/g•°C for all of your solutions.

4. Calculate the energy per gram of solute (in J/g) for each solution.

5. Using the periodic table, calculate the molar mass of each of the solutes.

6. Calculate the molar heat of solution for each solute from the formula molar heat of solution = (energy/gram) × (gram/mole of solute).

7. Your teacher will give you the accepted values of the molar heat of solution for each of your solutes. Use those values to calculate the percent error of your experimental value, using the following equation.

$$\%\ \text{error} = \frac{|\text{accepted value} - \text{experimental value}|}{\text{accepted value}} \times 100$$

Critical Thinking: Analysis and Conclusions

1. When sodium chloride dissolves in water, the ions dissociate. The equation for this reaction is $NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$. Write similar ionic equations to show the dissociation in water of each of the solutes used in this investigation. *(Making inferences)*

Name _____

Lab 34

2. Which reactions are exothermic? Which are endothermic? *(Classifying)*

3. According to your answer to the previous question, assign the proper sign to each of the molar heats of solution you calculated and write the ΔH for each reaction. *(Classifying)*

4. Rewrite each of the ionic equations from Question 1 showing the molar heat of solution as a reactant or a product. *(Making inferences)*

Critical Thinking: Applications

1. Suggest two uses for these solution reactions in sports injuries or camping. *(Applying concepts)*

2. Which solid(s) from this investigation could be used in each of your answers in the previous question? Explain your reasoning. *(Giving examples)*

3. What other kinds of considerations should you take into account if your suggestions were to be put into use? *(Applying concepts)*

4. How could you reduce the experimental errors in this investigation? Explain your reasoning. *(Designing experiments)*

Accepting Truth by Indirect Evidence

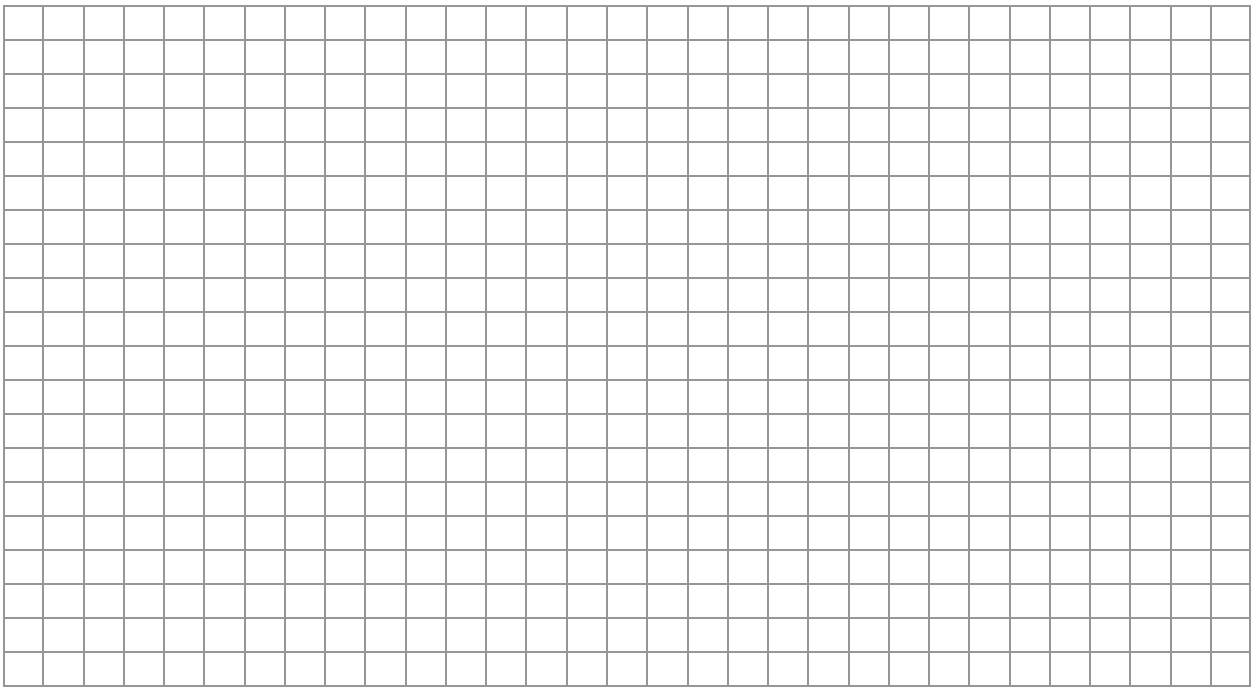
Problem: Using two different methods, what is the mass of the cup? (Do not take the pennies out of the cups!)

Equipment: 6 cups, each with 1 to 6 pennies
Balance
Ruler

Procedure:
1.
2.
3.
4.
5.
6.
7.
8.
9.
10.

Calculations:

Graphing:



Isotopes and Atomic Mass

Objectives

- Determine the average weight of each isotope of the fictitious element vegium.
- Determine the relative abundance of isotopes of vegium.
- Calculate from experimental data the atomic mass of vegium.

Introduction

Isotopes are atoms of the same atomic number having different masses due to different numbers of neutrons. The atomic mass of an element is the weighted average of the masses of the isotopes of that element. The weighted average takes into account both the mass and relative abundance of each isotope as it occurs in nature. The relative abundance and masses of small atomic particles are measured in the laboratory by an instrument called a mass spectrometer (Section 4D) The mass spectrometer separates particles by mass and measures the mass and relative abundance of each. From these data a weighed average is calculated to determine the atomic mass of the element.

Purpose

In this lab you will carry out experiments and perform the necessary calculations to determine the atomic mass of the fictitious element vegium. The three different isotopes of vegium are red beanium, black beanium, white beanium, brown beanium. As in real elements, these isotopes are collections of particles having different masses. Your job will be to obtain a sample of vegium and determine the relative abundance of each isotope and the mass of each type of particle. From this data you will calculate the weighted average mass, or atomic mass of vegium. Unlike real isotopes, the individual isotopic particles of vegium differ slightly in mass, so you will determine the average mass of each type of isotopic particle. Then you can calculate the weighted average mass, or "atomic mass" of vegium.

Safety

- Behave in way that is consistent with a safe laboratory environment.

Equipment

- a sample of vegium in a plastic cup
- small-scale balance

Procedure

Carry out the following steps, and record your results in Table 5.1

1. Weigh all the beans, all the peas, and all the corn.
2. Count all the beans, all the peas, all the corn.
3. Divide the mass of each isotope (bean, peas, corn) by the number of each isotope to get the average mass of each isotope.
4. Divide the number of each isotope by the total number of all of the isotopes to get the relative abundance of each isotope.
5. Multiply the relative abundance from Step 4 by 100 to get the percent abundance of each isotope
6. Multiply the relative abundance from Step 4 by the average mass of each isotope to get the relative weight of each isotope.
7. Add the relative weights to get the average mass of all particles in vegium, the "atomic mass."

Note: When you weigh the various types of vegetables, you may encounter some problems. For example, the sample of beans might be too large to weigh on your balance. You might solve this problem by making more weights or by using a larger counterweight on your balance. This approach increases your balance's capacity. Keep in mind that it also results in a heavier beam, which reduces the sensitive of your balance.

Alternatively, you might weigh a portion of your vegetables, say half, and then multiply your result by two (or a fifth and multiply by five). The beans are of various sizes, so if you weigh just one bean and multiply by the number of beans to get the total weight of beans, a significant error might result. Weigh a large enough sample so you get a good estimate of the average weight of a bean.

Data

Table 5-1 "Atomic Mass" of Vegium

	Black beans	Split-Peas	Corn		Total
1. Mass of isotope					
2. Number of atoms					
3. Average mass isotope					
4 Relative abundance					
5. Percent of isotope					
6. Relative weight					

Cleaning up

Place the entire sample of vegium back in the plastic cup. Make sure that none of the particles are in the sink or on the floor.

Questions for Analysis

Use your Experimental Data and what you learned in this lab to answer the following questions.

1. Which of your data in Table 5-1 must be measured and which can be calculated?

2. In all except Step 3 in Table 5-1 the numbers in the total column can be obtained by adding the numbers across each row. Step 3 is an exception because it does not take into account the fact that there are different numbers of each kind of particle. Rather than add across, calculate this number in the same way you calculated the other numbers in row 3.

3. What is the difference between percent and relative abundance?

4. What is the result when you total the individual percentages? The individual relative abundance?

5. The percentage of each vegetable tells you how many of each kind of vegetable there are in every 100 particles. What does relative abundance tell you?

6. Compare the total values for Steps 3 and 6 on Table 5-1.

7. Why can't atomic masses be calculated the way the total for row 3 is calculated?

8. Explain any differences between the atomic mass of your vegium sample and that of your neighbor. Explain why the difference would be smaller if larger samples were used.

Practice WKST
Electron Configuration
Using the Aufbau Diagram
and a List of Elements

Name_____
Date_____ Block____

Background

The electrons in an atom occupy distinct ***principal energy levels***. To be located in any each of these principal energy levels, electrons must have the required energy for the level. The principal energy levels are numbered **1, 2, 3, 4, 5, 6,** and **7** for the atoms of the known elements.

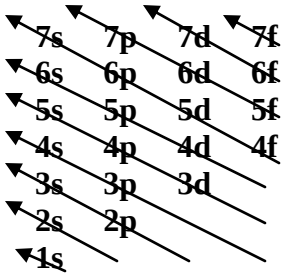
Each principal energy level is subdivided into ***sublevels***. Each sublevel has its own energy requirement for electrons. The sublevels are labeled ***s, p, d, and f.***

Within each sublevel there is a specific number of ***orbitals***. The orbitals within a given sublevel have the same energy requirements for electrons. Each orbital can "hold" a maximum of two (2) electrons.

Rules for "filling" energy levels:

1. **Electrons enter the lowest principal energy level available.**
2. **Within a principal energy level: electrons enter the lowest energy sublevel available.**
3. **Within a sublevel: electrons enter each orbital one electron at a time until all orbitals have one electron. Then additional electrons enter each orbital until 2 electrons are in each orbital. Once all orbitals in a sublevel are filled (each with 2 electrons), the next electron enters the next higher energy sublevel.**

The Aufbau diagram below illustrates the order of filling orbitals and sublevels. Start at the 1s sublevel and follow the arrows. For example: 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, etc.



Objectives for this worksheet:
To write the electron configurations for many elements.
To relate the electron configurations of the elements to their chemical behavior (whether the atom tends to lose or gain electrons).

MATERIALS
Per student: **4** copies of the electron configuration patterns (4 elements per sheet)
copy of the periodic table of elements
list of ions (atoms that have lost or gained electrons and have become charged particles; atoms that lose electrons become positively charged; atoms that gain electrons become negatively charged)

PROCEDURE
1. Each student will be assigned between 12 and 16 elements.
2. For each of the elements you are assigned, you must complete one electron configuration pattern.
3. With the other members of your group, arrange the electron configuration grids according to increasing atomic number.

SAVE THESE ELECTRON CONFIGURATION SHEETS FOR USE IN A LAB ON THE PERIODIC TABLE. You will use these sheets to answer Analysis and Conclusion items to relate electron configuration with the placements of the elements in the periodic table.

Background information: **Ions** form when neutral atoms gain or lose their outermost electrons. These outermost electrons are called the **valence electrons** and are responsible for the chemical behavior of the atoms.

Generally:
a) atoms having 1, 2, or 3 valence electrons tends to lose all of their valence electrons and form ions having a positive electrical charge. The size of the charge is equal to the number of electrons that have been lost. For example, a lithium atom has 1 valence electron; when that electron is lost, a lithium ion is formed having a +1 charge;

b) atoms having 5, 6, or 7 valence electrons tend to gain enough additional electrons to have a total of 8 valence electrons and will form ions having a negative electrical charge. The size of the charge is equal to the number of electrons that have been gained. For example, a nitrogen atom has 5 valence electrons; when the nitrogen atom gains 3 additional electrons, a nitrogen ion is formed having 8 valence electrons and a -3 charge.

c) atoms having 4 valence electrons will lose or gain valence electrons depending on whether the element is a metal or nonmetal. We'll deal with these elements in the next part of this unit on the Periodic Table.

Using the background information given above, your electron configuration pattern sheets and the list of ions, complete the following table. The first has been done as an example.

Element Name	Element Symbol	Atomic Number	Number of Protons	Number of Electrons in the Neutral Atom	Number of Valence Electrons	Tendency to GAIN or LOSE Valence Electrons	Charge on the Ion (kind and size)	Complete Ion Symbol
hydrogen	**H**	**1**	**1**	**1**	**1**	**lose**	**+1**	**H^{+1}**
lithium								
beryllium								
nitrogen								
oxygen								
fluorine								
sodium								
magnesium								
aluminum								
phosphorus								
sulfur								
chlorine								
potassium								
calcium								
strontium								
zinc								
barium								

Atomic number: _____ Element name: _______________ Element symbol: ______ Number of electrons: ______

1 *s*	2 *s*	2 *p*	3*s*	3 *p*	3 *d*	4 *s*	4 *p*	4 *d*	4 *f*
○	○	○○○	○	○○○	○○○○○	○	○○○	○○○○○	○○○○○○○

Atomic number: ______ Element name: _______________ Element symbol: ______ Number of electrons: ______

1 *s*	2 *s*	2 *p*	3*s*	3 *p*	3 *d*	4 *s*	4 *p*	4 *d*	4 *f*
○	○	○○○	○	○○○	○○○○○	○	○○○	○○○○○	○○○○○○○

Atomic number: ______ Element name: _______________ Element symbol: ______ Number of electrons: ______

1 *s*	2 *s*	2 *p*	3*s*	3 *p*	3 *d*	4 *s*	4 *p*	4 *d*	4 *f*
○	○	○○○	○	○○○	○○○○○	○	○○○	○○○○○	○○○○○○○

Atomic number: ______ Element name: _______________ Element symbol: ______ Number of electrons: ______

1 *s*	2 *s*	2 *p*	3*s*	3 *p*	3 *d*	4 *s*	4 *p*	4 *d*	4*f*
○	○	○○○	○	○○○	○○○○○	○	○○○	○○○○○	○○○○○○○

Atomic number: _____ Element name: _______________ Element symbol: ______ Number of electrons: ______

○	○	○○○	○	○○○	○○○○○	○	○○○	○○○○○	○○○○○○○
1 *s*	2 *s*	2 *p*	3*s*	3 *p*	3 *d*	4 *s*	4 *p*	4 *d*	4 *f*

Atomic number: ______ Element name: _______________ Element symbol: ______ Number of electrons: ______

○	○	○○○	○	○○○	○○○○○	○	○○○	○○○○○	○○○○○○○
1 *s*	2 *s*	2 *p*	3*s*	3 *p*	3 *d*	4 *s*	4 *p*	4 *d*	4 *f*

Atomic number: ______ Element name: _______________ Element symbol: ______ Number of electrons: ______

○	○	○○○	○	○○○	○○○○○	○	○○○	○○○○○	○○○○○○○
1 *s*	2 *s*	2 *p*	3*s*	3 *p*	3 *d*	4 *s*	4 *p*	4 *d*	4 *f*

Atomic number: ______ Element name: _______________ Element symbol: ______ Number of electrons: ______

○	○	○○○	○	○○○	○○○○○	○	○○○	○○○○○	○○○○○○○
1 *s*	2 *s*	2 *p*	3*s*	3 *p*	3 *d*	4 *s*	4 *p*	4 *d*	4 *f*

Atomic number: ______ Element name: _______________ Element symbol: ______ Number of electrons: ______

1 *s*	2 *s*	2 *p*	3*s*	3 *p*	3 *d*	4 *s*	4 *p*	4 *d*	4 *f*
○	○	○ ○ ○	○	○ ○ ○	○ ○ ○ ○ ○	○	○ ○ ○	○ ○ ○ ○ ○	○ ○ ○ ○ ○ ○ ○

Atomic number: ______ Element name: _______________ Element symbol: ______ Number of electrons: ______

1 *s*	2 *s*	2 *p*	3*s*	3 *p*	3 *d*	4 *s*	4 *p*	4 *d*	4 *f*
○	○	○ ○ ○	○	○ ○ ○	○ ○ ○ ○ ○	○	○ ○ ○	○ ○ ○ ○ ○	○ ○ ○ ○ ○ ○ ○

Atomic number: ______ Element name: _______________ Element symbol: ______ Number of electrons: ______

1 *s*	2 *s*	2 *p*	3*s*	3 *p*	3 *d*	4 *s*	4 *p*	4 *d*	4 *f*
○	○	○ ○ ○	○	○ ○ ○	○ ○ ○ ○ ○	○	○ ○ ○	○ ○ ○ ○ ○	○ ○ ○ ○ ○ ○ ○

Atomic number: ______ Element name: _______________ Element symbol: ______ Number of electrons: ______

1 *s*	2 *s*	2 *p*	3*s*	3 *p*	3 *d*	4 *s*	4 *p*	4 *d*	4*f*
○	○	○ ○ ○	○	○ ○ ○	○ ○ ○ ○ ○	○	○ ○ ○	○ ○ ○ ○ ○	○ ○ ○ ○ ○ ○ ○

Atomic number: _____ Element name: _______________ Element symbol: ______ Number of electrons: ______

○	○	○○○	○	○○○	○○○○○	○	○○○	○○○○○	○○○○○○○
1 *s*	2 *s*	2 *p*	3*s*	3 *p*	3 *d*	4 *s*	4 *p*	4 *d*	4 *f*

Atomic number: ______ Element name: _______________ Element symbol: ______ Number of electrons: ______

○	○	○○○	○	○○○	○○○○○	○	○○○	○○○○○	○○○○○○○
1 *s*	2 *s*	2 *p*	3*s*	3 *p*	3 *d*	4 *s*	4 *p*	4 *d*	4 *f*

Atomic number: ______ Element name: _______________ Element symbol: ______ Number of electrons: ______

○	○	○○○	○	○○○	○○○○○	○	○○○	○○○○○	○○○○○○○
1 *s*	2 *s*	2 *p*	3*s*	3 *p*	3 *d*	4 *s*	4 *p*	4 *d*	4 *f*

Atomic number: ______ Element name: _______________ Element symbol: ______ Number of electrons: ______

○	○	○○○	○	○○○	○○○○○	○	○○○	○○○○○	○○○○○○○
1 *s*	2 *s*	2 *p*	3*s*	3 *p*	3 *d*	4 *s*	4 *p*	4 *d*	4 *f*

Flame Test Lab

When elements are heated to high temperatures, they may enter an excited state. In an excited state, the electrons move to higher energy levels. The change in energy that occurs when the excited electrons return to their **ground** state causes the substance to emit light. The observed colors or spectrum of the substance is caused by the visible wavelengths of the light emitted. Since each element emits a unique set of wavelengths, emission (line) spectra can be used as a tool to identify the elements. One method used to demonstrate the emission spectrum of a substance is the flame test. Using this method, a small amount of a substance is heated and the characteristic glow of the substance is observed. In this experiment you will perform a flame test on several metallic salts. Based on your observation, you will develop a reference table which lists the flame color for each metal ion. You will then perform a flame test on an unknown substance. By comparing your observations to the data in your reference table, you will be able to identify the metal ion in the unknown substance. Finally, you will use cobalt glass as a tool for identifying the components of a metallic salt mixture.

Purpose: Determine the purpose from the information above.

Hypothesis: Predict how successful you will be in determining your unknown

Pre-lab Questions:

1. List any safety concerns in this experiment.
2. Describe the process of energy absorption and emission of light by an atom. (How do we get line spectra?)
3. Why are emission and absorption spectra important? (How can they be used?)
4. What is meant by the ground state of an electron?

Materials: Goggles and apron Bunsen burner
Co glass squares Beaker
Wood splints soaked in known and unknown solutions

Procedure: You will work in your lab group for this activity.

1. Obtain wood splints soaked in each of the metal ion solutions listed below.
2. Watch the color of the flame as the wooden splint is burned.
3. Record your observations.
4. After completing all known solutions, obtain the unknowns.

5. Repeat steps 2 and 3 for the unknowns.
6. Using the information on the known solutions, identify each unknown and record in your table.
7. Clean up as instructed by your teacher

Data:

Construct a data table like the one below

Metal Ion	**Flame Color**	**Unknowns**	**Flame Color**	**Identity**
Sodium, Na^{+}		#1		
Barium, Ba^{+2}		#2		
Calcium, Ca^{+2}		#3		
Copper, Cu^{+2}		#4		
Potassium, K^{+}		#5		
Strontium, Sr^{+2}		#6		
Lithium, Li^{+}		#7		
Iron, Fe^{+3}		#8		
Na^{+} and K^{+}				
Co glass, Na^{+}, K^{+}				
Na^{+} and Co glass				
K^{+} and Co glass				

Analysis Questions:

1. Explain how a flame test can be used to identify an element.
2. Explain the difference between emission and absorption spectra.
3. Based on your results and observations, would this method be practical to determine metals in a mixture? Why or why not?
4. Explain the reason potassium was visible when using the cobalt glass.
5. Give at least two reasons why the flame test is sometimes invalid.
6. For each ion, state how many valence electrons have been removed.
7. What were the approximate wavelengths of the light that was emitted for each of the elements?

Conclusion

Comments

Flame Test Colors

Symbol	Element	Color
As	Arsenic	Blue
B	Boron	Bright green
Ba	Barium	Pale/Yellowish Green
Ca	Calcium	Orange to red
Cs	Cesium	Blue
Cu(I	Copper(I)	Blue
Cu(II)	Copper(II) non-halide	Green
Cu(II)	Copper(II) halide	Blue-green
Fe	Iron	Gold
In	Indium	Blue
K	Potassium	Lilac to red
Li	Lithium	Magenta to carmine
Mg	Magnesium	Bright white
Mn(II)	Manganese(II)	Yellowish green
Mo	Molybdenum	Yellowish green
Na	Sodium	Intense yellow
P	Phosphorus	Pale bluish green
Pb	Lead	Blue
Rb	Rubidium	Red to purple-red
Sb	Antimony	Pale green
Se	Selenium	Azure blue
Sr	Strontium	Crimson
Te	Tellurium	Pale green
Tl	Thallium	Pure green
Zn	Zinc	Bluish green to whitish green

Ionic or Covalent Lab

Introduction:
Compounds can be classified by the types of bonds that hold their atoms together. Ions are held together by ionic bonds in ionic compounds; atoms are held together by covalent bonds in covalent compounds.
You cannot tell whether a compound is ionic or molecular simply by looking at a sample of it because both types of compounds can look similar. However, simple tests can be done to classify compounds by type because each type has a set of characteristic properties shared by most members. Ionic compounds are usually hard, brittle, water-soluble, have high melting points, and can conduct electricity when dissolved in water. Covalent compounds can be soft, hard, or flexible; are usually less water-soluble; have lower melting points; and cannot conduct electricity when dissolved in water.

Problem:
How can you identify ionic and covalent compounds by their properties.

Objectives:
- Examine the properties of several common substances.
- Interpret the property data to classify each substance as ionic or molecular.

Materials:
- Glass microscope slide
- Grease pencil
- Hot plate
- Spatula
- 4 small beakers (50 or 100 mL)
- Stirring rod
- Balance
- Conductivity tester
- Small graduated cylinder
- Thermometer (up to 150°C)
- 0.5 g samples of unknowns

Safety Precautions:
- Burns …
- Electrical …
- Chemical …
- Cuts …

Never use a grease pencil to write on the white label area of a beaker!!!

Procedure:

A
B
C
D

1. Use a grease pencil to draw lines dividing a glass slide into four parts. Label the parts A, B, C, and D.
2. Record all data in the table under Data and Observations.
3. Use a spatula to place a small amount of the first of the four substances on section A of the slide.
4. Repeat step 3 with your other three substances on sections B, C, and D. Be sure to use a clean spatula for each sample.
5. Place the slide on a hot plate. Turn the heat setting to medium and begin to heat the slide.
6. Gently hold a thermometer so that the bulb just rests on the slide. Be careful not to disturb your compounds.
7. Continue heating until the temperature reaches 90°C. Observe each section on the slide and record which substances have melted and the temperature at which they melted. Turn off the hot plate. Dispose of the microscope slide. Put away any equipment.
8. Wash and rinse 4 beakers with distilled water, label each beaker (E, F, & G) with a pencil.

Ionic and Covalent Lab

Name _____
Period ____ Date _____

9. Add 0.5 g of compound E to one beaker and add 10 mL of DI water. Test the conductivity of controls E, F, & G. Record data in chart below.
10. Clean, rinse and reuse the beakers to test unknowns compounds A – D. Erase pencil label and re-label on the white area.
11. Weigh 0.5 g of each of the 4 substances and place the weighed samples in their beakers.
12. Add 10 mL of distilled water to each beaker.
13. Stir each substance, using a clean stirring rod for each sample. Note on the table below whether or not the sample dissolved completely.
14. Test each substance for the presence of electrolytes (ions) by using a conductivity tester. Record below whether or not each acts as a conductor.

Analyze and Conclude

1. Classifying: Complete your data table below by classifying each of the substances you tested as ionic or covalent compounds based on your observations.

Data and Observations

Substance	Did it melt? (Y/N)	Melting Point Temperature (°C)	Did it dissolve in water? (Y/N)	Did the solution conduct electricity? (Y/N)	Classification
Unknown A					
Unknown B					
Unknown C					
Unknown D					
E NaCl (control)	**XXXX**	**XXXX**	**XXXX**		
F DI H_2O (control)	**XXXX**	**XXXX**	**XXXX**		
G Tap H_20 (control)	**XXXX**	**XXXX**	**XXXX**		

Apply and Assess:

1. What are the differences in properties between ionic and covalent compounds?

2. How did the melting points of the ionic compounds and the covalent compounds compare?

3. How did the conductivity of the ionic and the covalent compounds compare?

4. What happened to the bonds between the molecules when a substance melted?

Chemistry Tie Dye Lab

Purpose: To study the chemistry of dyes and dying

Background: Some dyes only stain the cloth, and wash out a little each time the cloth is laundered. High Quality dyes (fiber-active dyes) actually chemically (covalent) bond to the molecules of the fabric and can never be washed out. The dye molecules carry a "chromophore", which absorb varying spectrums of light, allowing only certain spectrums to reflect.

Cotton is made of long strands of cellulose molecules, all twisted together. Cotton is ideal for dying because the fibers are naturally hollow, and the dye molecules will form bonds on both the inside and outside of the fiber. If you put molecules of dye and cotton together, very little will happen until the atoms on the surfaces of the molecules are prepared for bonding. This can be done by either increasing the temperature or the pH of the fiber and dye. In this lab, we will do the latter, increasing the pH by soaking the fabric in calcium carbonate. This releases a H atom from the cellulose molecule, and a Cl atom from the dye, making both suitable for bonding.

The dye is allowed to react in a desirable host environment for up to 24 hours. After this time, the bonding sites on the cellulose should be saturated with dye molecules. Excess dye molecules that have not bonded permanently are washed away using warm water rinse and a dye-carrying detergent

Procedure: The whole point of tie dying is to prevent the dye from reaching the fabric evenly. Any place that the dye can't reach will stay white, or a lighter color. You can accomplish this by folding the fabric, tying it with string, using rubber bands, etc.

Step 1: With a permanent marker, write your name on the cloth you will be tie dying. There will be over 150 clothing items in the classroom during this lab; ***be sure yours is marked to avoid losing it***.

Step 2: Soak your cloth in the calcium carbonate solution overnight. Be sure you put it in the tub marked for your period.

Step 3: on the following day, put on a pair of gloves, remove your cloth from the carbonate soaking tub and wring out. Using the below diagrams, tie you cloth in one (or a combination of) the methods shown using string or rubber bands.

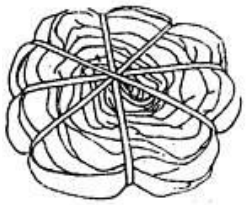
Step 4: Dye application. Apply the dye using the applicator bottles. *Be aware that mixing will occur where the dyes come in contact with each other*. Using complimentary colors (purple and yellows, blues and orange, or red and green) near each other usually produces a brownish black color. Also, the more dye you put on a given spot, the less white will remain on the final product.

Step 5: Place your dye-soaked item in a grocery or other plastic bag marked clearly with your name. No pooling of liquid should appear in the bag.

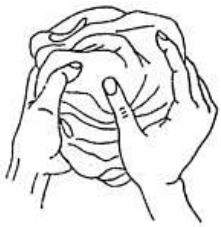
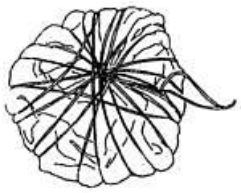
Step 6: After the cloth has soaked for at least 24 hours, open the bag, remove the item, and rinse it several times, before removing strings or rubber bands. Once it is rinsing clean, remove the ties and rinse several times again until it is rinsing clean. At this point you can hang it to dry or place it in another bag to take home and wash.

The first time you wash it, place it alone in the washing machine with just a small squirt of dishwashing liquid like Dawn or Joy. **Do not use laundry detergent**.

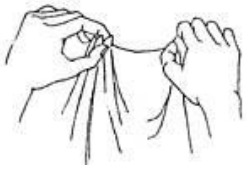
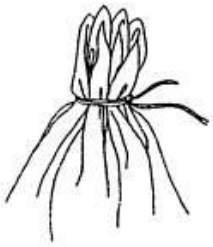

Knots

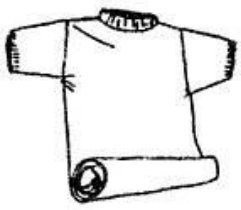
Spirals – lay material flat. With thumb and forefinger, start twisting the material at what will be the center of the spiral. Continue twisting and pushing down on the material until it is wound flat like a pancake. Apply rubber bands to hold in place, with bands intersecting in the center. Apply dye, or place in the dye bath.

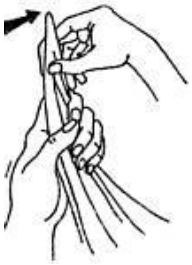
Electric Bunching – gather cloth together in small bunches until shaped like a ball. Try to expose as much surface as possible. Wrap strings or rubber bands around the ball to retain its shape. Set gently in the dye bath, turning occasionally.

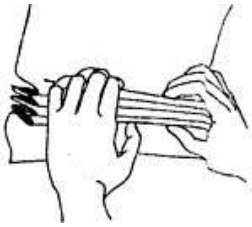
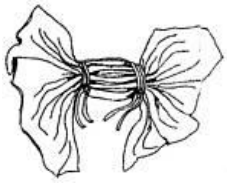
Rosettes are many small circles touching or overlapping each other. Make a few dots on the cloth in a pattern. Pick up each dot and transfer to the other hand. Wrap rubber bands around the base of the circles.




Stripes - roll the cloth very loosely, into a long tube. The stripes will be at right angles to the tube. Tie at intervals as far apart as you want the stripes. Make sure knots or rubber bands are tight.


Circles – pick up the cloth with thumb and forefinger at the center of the circle. With the other hand, try to make neat and even pleats around the circle. Smooth the cloth, hold tightly at the base, and let go of the top. With string or rubber band, secure tightly at base. Continue wrapping to the tip and back. Make sure ties are very tight.

Pleats - Lay cloth on flat surface. Position fingers about an inch or two in front of your thumbs, and pinch the fabric to raise a fold. Continue to pinch up until you reach the end of the cloth. You can change directions as often as you want by gathering more material in one hand than in the other. Loop rubber bands or string very tightly around all the pleats several times and knot. You can use as many ties as you want.

Name _____ Date _____ Class _____

Models of Molecular Compounds

Lab 22

Text reference: ***Chapter 8***

Introduction

Why should people care about the shapes of molecules? Consider that the properties of molecules, including their role in nature, depend not only on their molecular composition and structure, but their shape as well. Molecular shape determines a compound's boiling point, freezing point, viscosity, and the nature of its reactions.

The geometry of a small molecule can be predicted by examining the central atom and identifying the number of atoms bonded to it and the number of unshared electron pairs surrounding it. The shapes of molecules may be predicted using the VSEPR rule, which states that electron pairs around a central atom will position themselves to allow for the maximum amount of space between them.

Covalent bonds can be classified by comparing the difference in electronegativities of the two bonded atoms. If the difference in electronegativities is less than or equal to 0.4, the bond is called a nonpolar covalent bond. If the difference in electronegativities is between 0.5 and 1.9, a polar covalent bond exists. (If the difference in electronegativities is greater than 2.0, an ionic bond results.) In a polar covalent bond, the electrons are more attracted to the atom with the greater electronegativity, resulting in a partial negative charge on that atom. The atom with the smaller electronegativity value acquires a partial positive charge.

Molecules made up of covalently bonded atoms can be either polar or nonpolar. The geometry of the molecule determines whether it is polar or not. For example, if polar bonds are symmetrically arranged around a central atom, their charges may cancel each other out and the molecule would be nonpolar. If, on the other hand, the arrangement of the polar bonds is asymmetrical, the electrons will be attracted more to one end of the molecule and a polar molecule or dipole will result.

Ball-and-stick models can be used to demonstrate the shapes of molecules. In this experiment, you will construct models of covalent molecules and predict the geometry and polarity of each molecule.

Pre-Lab Discussion

Read the entire laboratory investigation and the relevant pages of your textbook. Then answer the questions that follow.

1. What is a covalent bond? _____

_____

2. What is a dipole? _____

_____

Name _____

3. What two factors determine whether a molecule is polar or not?

_____

_____

4. List the five different molecular geometries that you will be studying in this investigation. _____

_____

5. Calculate the electronegativity difference and predict the type of bond for the following examples: (Refer to Figure 7–19 in your text for a list of electronegativities.)

   a. Na—Cl _____

   b. C—H _____

   c. S—O _____

   d. N—N _____

Problem

How can the polarity of molecules be predicted from their geometry and the types of bonds they contain?

Materials

safety goggles
ball-and-stick model set

Safety 

Wear your goggles at all times during the investigation. Note the caution alert symbols here and with certain steps of the Procedure. Refer to page *xi* for the specific precautions associated with each symbol.

Procedure

1. Put on your goggles. Construct ball-and-stick models of the following compounds:

H_2	HBr	H_2O
PH_3	CH_4	HClO
N_2	CH_3NH_2	CH_3Cl
H_2CO	C_2H_2	H_2O_2
HCOOH	HCN	

2. For each of the preceding compounds, complete the Data Table in the Observations section. As an example, the first line of the Data Table has been filled in for you.

3. When you have completed this investigation, take apart your models and return the model set to your teacher. Clean up your work area and wash your hands before leaving the laboratory.

Name _____

Lab 22

Observations

DATA TABLE Structure and Polarity of Molecules

Formula	Electron Dot Structure (Lewis)	Structural Formula	Shape of Molecule	Molecular Polarity
H_2	H:H	H—H	Linear	Nonpolar
HBr				
H_2O				
PH_3				
CH_4				
HClO				
N_2				
CH_3NH_2				
H_2CO				
C_2H_2				
CH_3Cl				
HCOOH				
HCN				
H_2O_2				

Lab 22

Name _____

Critical Thinking: Analysis and Conclusions

1. Explain how you used the molecular shapes to predict molecular polarity. Support your answer with examples from the results of this investigation. (*Classifying*) _____

2. List the advantages and disadvantages of using ball-and-stick models to construct molecules. (*Developing models*) _____

Critical Thinking: Applications

1. Based on your results, predict the type of bonding, molecular geometry, and molecular polarity of the following molecules. (*Making predictions*)
   a. HI _____
   b. SH_2 _____
   c. NH_3 _____
   d. CO_2 _____

2. The polarity of a substance can have a great effect on its reactivity and solubility. A rough rule of thumb for solubility is "like dissolves like." Knowing this general rule, what can you predict about the polarity of alcohol if you know that alcohol dissolves in water? Why do you think that water is not used to dissolve greasy stains and dirt at dry cleaners? (*Applying concepts*) _____

Going Further

1. Use balloons to create three-dimensional models of the five different molecular geometries discussed in this investigation.
2. Research what is meant by the term *isomer*. Give examples of molecular isomers.

Sweetly Balanced Equations

Using Simulations with Candy to Help Understand the Balancing of Chemical Equations

Introduction

Pieces of candy will be used to represent atoms in chemical equations. Different colors will represent different atoms. Conservation of atoms in a chemical equation will be shown by having the same number and kinds of atoms on each side of the equation. One lab partner will use his/her candy to simulate the reactant (left) side of the equation, and the other partner will use his/her candy to simulate the product (right) side.

Materials (per person):

- candy (see kinds and amounts in table on next page)
- small square of paper with **=** printed on it (1 per group)
- small square of paper with **+** printed on it (2 per group)
- clean paper napkin
- clean piece of copier paper

IMPORTANT NOTES:

• The candy used in this activity is edible, and it is important that cleanliness and common-sense sanitary precautions be observed when handling it.

• Wash your hands, since you will be handling candy you may eventually want to eat.

• Follow class instructions for obtaining candy in supply containers.

• Place candy on the napkin or the copier paper to keep it clean.

• ONCE YOU HAVE HANDLED CANDY, DO NOT RETURN IT TO A SUPPLY CONTAINER OR GIVE IT TO SOMEONE ELSE TO EAT; EITHER EAT IT YOURSELF OR THROW IT AWAY.

California Science Standards Addressed

• Grade 7; Physical Science; Reactions; 5b; Students know the idea of atoms explains the conservation of matter: **In chemical reactions the number of atoms stays the same no matter how they are arranged**, so their total mass stays the same.
• Grades 9-12; Chemistry; Conservation of Matter and Stoichiometry; 3a**; Students know how to describe chemical reactions by writing balanced equations.**

References/Acknowledgement

• P. W. Atkins, **Molecules**, W.H. Freeman, 1987; a unique chemistry classic, showing relation of molecular structures to everyday substances and phenomena; great molecular drawings
• Nicole Hays, West Bloomfield High School, MI; **Candy Chemistry**; a shorter version of this activity, with no drawings; will print as a pdf; use a Google search to locate it on the web (the URL is LONG!)
• Tom Frost, my longtime close friend and colleague at Foothill High School, Pleasanton, CA; use of molecular drawings in equations; use of candy and other food as teaching tools in chemistry; many facets of teaching chemistry and helping students to learn

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Procedure

1. Obtain a Fun Size bag of m&m's, plus the specified number of pieces of other candy shown in the table below. If your bag of m&m's does not have enough green, blue or yellow m&m's, get them from the reserve stockpile. The numbers shown below are the minimum for you to be able to do the equation balancing. You may eat the brown, red and orange m&m's whenever you wish, but save all the green, blue and yellow ones until the activity is finished. Once the activity is over, you may eat any of your remaining candy.

ELEMENT	COLOR	CANDY	QUANTITY
H	white	miniature marshmallows	6
Cl	green	m&m's	4
O	red	Red Vine pieces (red)	7
N	blue	m&m's	3
C	black	Red Vine pieces (black)	3
Na	yellow	m&m's	2
Fe	silver	Hershey's Kisses	2

2. For equations (1) - (5) below, complete the following steps:

a. Try to balance the equation.

b. One of the two lab partners should use his/her pieces of candy to simulate the left side of the balanced equation on a piece of copier paper, and the other person should simulate the right side on a separate piece of paper. Use the small pieces of paper with **+** or **=** as appropriate. Make sure that there are the same number of pieces of each kind and color on each side of the equation.

c. When you and your lab partner have completed an equation, have your instructor check the balanced equation and the candy arrangement to verify that everything is correct. Your instructor will then initial in the space provided so that you will get credit.

(1) ___ Na + ___ Cl_2 = ___ NaCl initials _____

(2) ___ Na + ___ H_2O = ___ NaOH + ___ H_2 initials _____

(3) ___ CO + ___ NO = ___ CO_2 + ___ N_2 initials _____

(4) ___ Fe_2O_3 + ___ CO = ___ Fe + ___ CO_2 initials _____

(5) ___ C + ___ Fe_2O_3 = ___ CO + ___ Fe initials _____

4. The drawings for equations (6) - (10) below represent **unbalanced** chemical equations. For these equations, first use the drawings and the key provided to write the **unbalanced** equation, and then follow the same procedure for balancing, simulating with candy, and having your results initialed that you used for equations (1) - (5) above.

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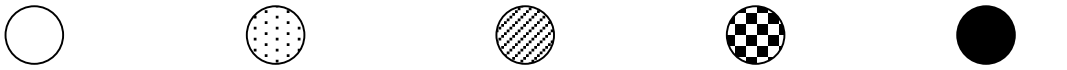
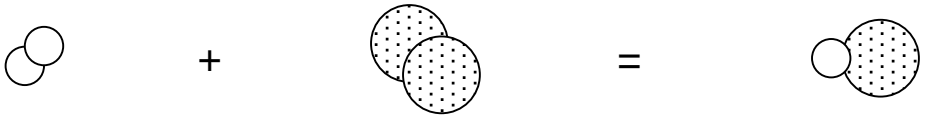
KEY:



HYDROGEN CHLORINE OXYGEN NITROGEN CARBON

(6)

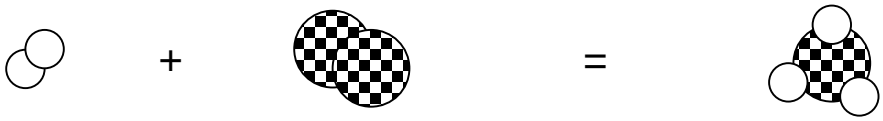


_____ initials_____

(7)

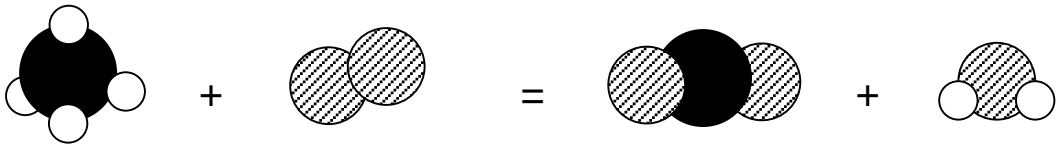
_____ initials_____

(8)



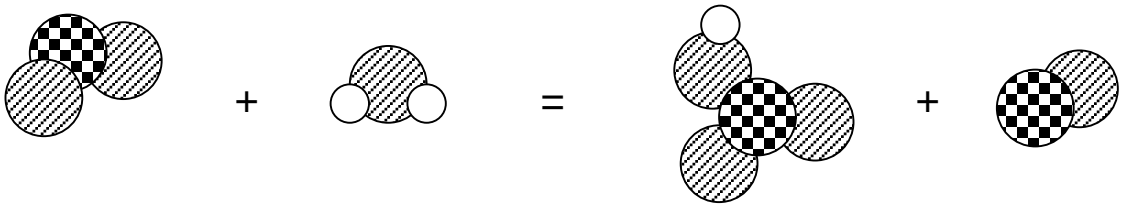
_____ initials_____

(9)



_____ initials_____

(10)



_____ initials_____

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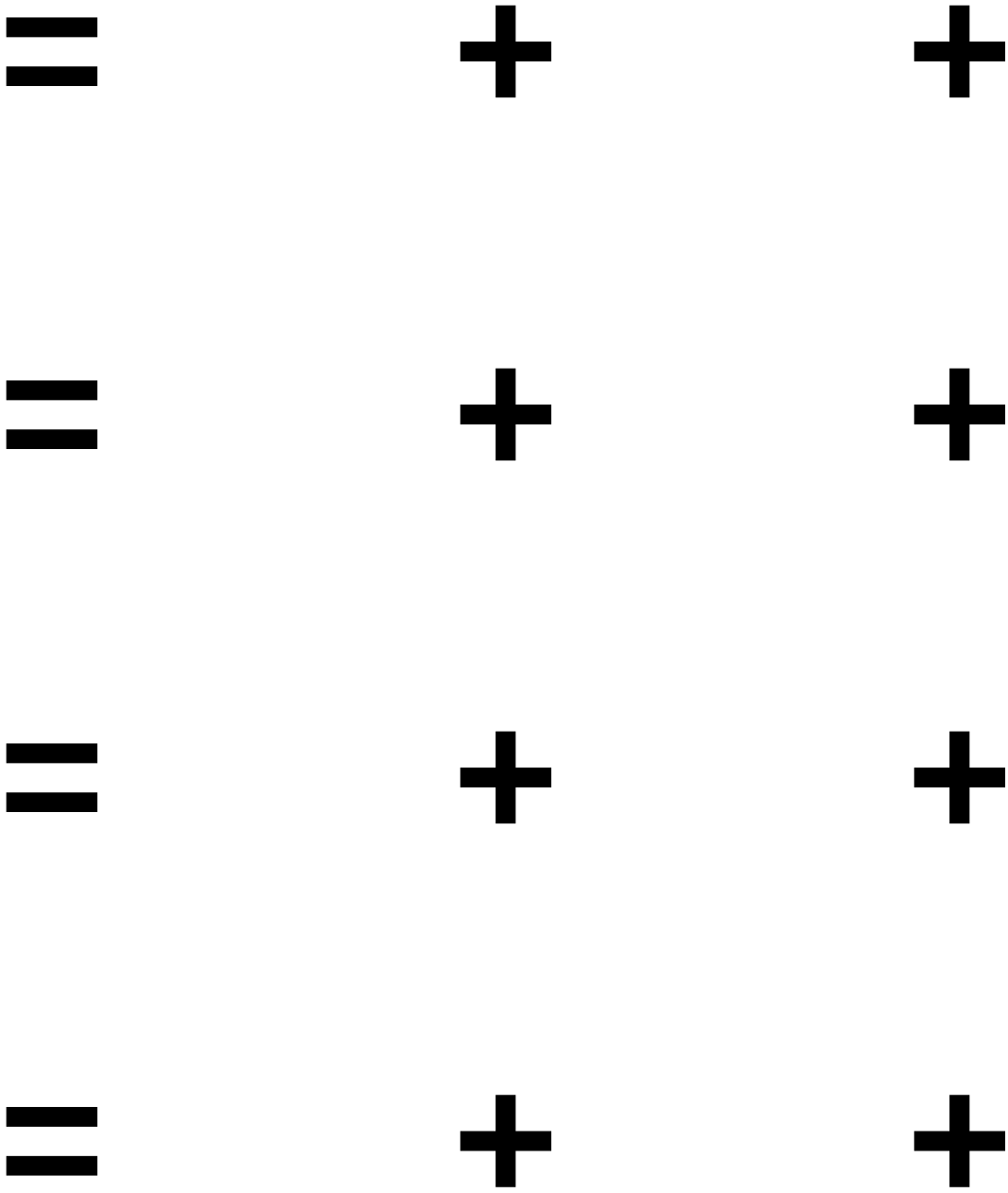
Cut out signs below so that each sign is on its own small square of paper. Each group requires one **=** sign and two **+** signs.



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Conservation of Mass – Part One

Purpose: To test how true the law of conservation of mass is, we will examine a simple chemical reaction and compare the mass of the system before and after the reaction.

Hypothesis: After reading the procedure for this lab, predict whether or not the law of conservation of mass will proven true. Why?

Pre-lab Questions:

1. State the Law of Conservation of Mass.

2. What are reactants?

3. What are products?

4. What are the signs you can look for that will tell you that a chemical reaction has a occurred?

Materials:

5% acetic acid solution (vinegar)
Sodium bicarbonate (baking soda)
Two clear plastic cups
Balance
Micro plunger
Graduated cylinder

Procedure:

1. Tap the micro plunger down into the baking soda, enough so that the bulb end is packed with a plug of the powder. (4-5 mL of baking soda should be enough to pack the bulb with a plug.)

2. Deliver the powder into one of the plastic cups. (Squeezing on the sides of the micro plunger can help loosen the plug.)

3. Use the graduated cylinder to measure out 100 mL of vinegar into the second plastic cup.

4. Place both the cups side by side on the balance to determine the total initial mass of the system. Record the mass in your data table.

5. Carefully pour the vinegar into the cup containing the baking soda, allowing vinegar to slowly run down the side of the cup's interior. Add just a little vinegar at a time to avoid having the reaction get out of control. Observe and record your observations about the reaction.

6. When the reaction has finished, place both cups back on the balance to determine the total final mass of the system. Calculate any change in mass, and record both in the data table.

7. Clean up your area and complete the analysis questions as a group.

Data:

Initial Mass (g)	
Final Mass (g)	
Change in Mass (g)	
Observations of Reaction	

Analysis:

1. What evidence was there that a chemical reaction occurred?

2. How did the final mass of the system compare with the initial mass of the system?

3. Considering your answer to #2, was the law of conservation of mass violated? How can you explain this?

Conservation of Mass – Part Two

Purpose: To redesign the Conservation of Mass lab to demonstrate that the Law of Conservation of Mass does hold true.

Materials:

See Part One
Hook-insert cap for bottle
2-L plastic soda bottle

Procedure:

1. You will need to write your procedure in your lab notebook. Be sure to include all necessary steps! No shortcuts!
2. Examine the plastic bottle and the hook-insert cap. Try to develop a modified procedure that will test the law of conservation of mass more accurately than Part One.
3. After you get your teacher to approve the procedure you designed, implement it. Record your data in a data table similar to the one you used in Part One.
4. If your procedure was successful, and your results reflect the conservation of mass, clean up and finish your lab.
5. If you were not successful, find a lab group that was successful, and discuss with them what they did and why they did it. Test their procedure yourself to determine whether the results of the other group are reproducible.
6. Be sure to clean up and dispose of all materials as directed by your teacher before leaving the lab area.

Data: use a chart similar to that from Part One

Analysis:

1. Was there any new evidence in Part II that a chemical reaction had occurred?
2. Identify the state of matter for each reactant in Part II.
3. Do your best to identify the state of matter for each product. *This may be tricky…*
4. Write a chemical reaction using words and/or chemical formulas describing the reaction you observed in this investigation.
5. What was different about the "systems" in Part I and Part II? Why was this important to your final outcome?

Conclusion:

Comments:

Reaction lab

Goal: To observe reactions and identify the type and products produced in those reactions and to compare reactions for their toxicity and use the 12 principles of green chemistry to examine critically the use of materials in a chemistry class.

Objectives: Students will …

- Perform various reactions
- Analyze these reactions and categorize them
- Analyze these reactions against the 12 principles of green chemistry

Materials: (per lab group)

- Goggles
- 3 test tubes
- steel wool
- copper (II) chloride solution
- rubber stopper
- crucible tongs
- calcium carbonate chips
- Copper (II) sulfate solution
- Potassium carbonate
- Graduated cylinder
- Catalayse/potato piece
- Dilute hydrochloric acid
- Zinc strip
- Bunsen Burner
- Copper wire
- Bunsen burner
- Flint lighter
- Crucible tongs
- Test tube rack
- Match
- Flint striker
- Hydrogen peroxide (5%)
- Sodium oxalate
- Calcium chloride
- Calcium Oxide
- 3 test tubes
- pH paper

Time Required: 1 x 60-75 minute class periods

Prerequisites: An understanding of the 12 principles of green chemistry

National Science Standards Met: S1, S2, S3, S6, S7, S8

Green Chemistry Principles Addressed: 1-12

Procedure:

- Explain to the students that today they will be exploring reactions through a series of labs.
- Chemists developing products or procedures in the lab are constantly having to evaluate reactions and decide which ones to use to meet a specific need and today the students will be evaluating reactions and looking at them with an eye for green chemistry.
- Hand out the pre-lab questions and ask students to complete the worksheet.
- Put students into lab groups of 2-3 and hand out the student lab sheet. Review the content and answer any questions.

Student Lab Sheet

Collect the following materials from the supply area and complete the lab procedures below:

Materials

- Goggles
- 3 test tubes
- steel wool
- copper (II) chloride solution
- rubber stopper
- crucible tongs
- calcium carbonate chips
- Copper (II) sulfate solution
- Potassium carbonate
- Graduated cylinder
- Catalayse/potato piece
- Dilute hydrochloric acid
- Zinc strip
- Bunsen Burner
- Copper wire
- Bunsen burner
- Flint lighter
- Crucible tongs
- Test tube rack
- Match
- Flint striker
- Hydrogen peroxide (5%)
- Sodium oxalate
- Calcium chloride
- Calcium Oxide
- 3 test tubes
- pH paper

Procedure:

Reaction type A

Part 1

1. Place a sample (10 mL) of copper (II) chloride solution in a test tube.
2. Take a piece of magnesium metal and sand it with the steel wool.
3. Record physical properties of both the solution and the magnesium metal.
4. Place the magnesium metal in the test tube. Record your observations. Complete Reaction type B and then return and observe the test tube again.

Part 2

1. Place a sample (10 mL) of dilute hydrochloric acid solution in a test tube.
2. Take a piece of zinc and sand it with the steel wool.
3. Record physical properties of both the solution and the magnesium metal.
4. Place the zinc metal in the test tube. Record your observations. Complete Reaction type B and then return and observe the test tube again

Reaction type B

Part 3

1. Obtain a sample (10 mL) of copper (II) sulfate. Place it in a test tube. Record its physical properties.
2. Take a sample (10 mL) of potassium carbonate. Record its physical properties.
3. Mix the two solutions in the test tube. Observe and record the outcome.
4. Leave the test tube in the test tube rack. Complete Reaction type C and then come back and look at the test tube to make sure your observations are complete

Part 4:

1 Obtain a sample (10 mL) of sodium oxalate. Place it in a test tube. Record its physical properties.
2 Take a sample (10 mL) of calcium chloride. Record its physical properties.
3 Mix the two solutions in one test tube. Observe and record the outcome.
4 Leave the test tube in the test tube rack. Complete Reaction type C and then come back and look at the test tube to make sure your observations are complete.

Reaction Type C

Part 5:

1. Take a sample (15 mL) of hydrogen peroxide (which is also known as: dihydrogen dioxide). Place it in a test tube.
2. Add a small sample of potatoe/catalayse into the test tube. Quickly place the rubber stopper LIGHTLYonto the test tube.
3. Observe what is happening. Allow the reaction to carry on for about 10 seconds.
4. Light a wooden splint using a match.. When the splint has burnt the a bit, blow out the flame. The splint should be glowing. Take the stopper off the test tube and place the glowing splint into the test tube.
5. Observe what happens to the splint.

Part 6:

1 Using a set of tongs obtain a sample of calcium carbonate.. Record its physical properties.
2 Place the calcium carbonate on a wire gauze (outside the clay circle if there is one present) and heat it in a blue Bunsen burner flame for 5 minutes. The temperature of the chip must reach 850 deg C.
3 Allow the chip to cool for 2 minutes. Inspect the chip and record its physical properties.

Reaction Type D

Part 7

1. Using a scoopula, obtain a small sample of Calcium Oxide and place into 2 test tubes. Record its physical properties
2. To one of the testtubes add 15 mL of water.
3. Add 15 mL of water to a 3rd test tube. .. Record its physical properties
4. Use pH paper to test the pH of each sample in each test tube (1. Calcium Oxide, 2 Calcium Oxide + water, and 3. Water)
5. Observe what happens to the pH of the water when it is reacted to the calcium oxide.

Part 8:

1. Take a small piece of copper wire and hold it at one end using the crucible tongs.. Record its physical properties.
2. Place the opposite end of the wire into the hottest part of the flame in the Bunsen burner (the blue part) for 30 s.
3. Remove the wire and examine it. After the wire is cooled, scrape the surface with the edge of a scoopula.

Student worksheet

Discussion Questions:

1. For each part listed, write a balanced chemical equation (NOTE: in part 5, the potato piece/catalayse is a CATALYST. It is not used up in the chemical reaction. Do not add it into the chemical equation for part 5) – *(4 points)*

   Reaction A_____

   Reaction B_____

   Reaction C_____

   Reaction D_____

2. For the products identified in your chemical equations, match the physical observations you made to the products predicted by your chemical equation (FOR EXAMPLE: If your reaction were to produce a yellow precipitate, then you would have to say: *In Part X, the yellow precipitate formed was lead (II) iodide*). Complete for all parts in the lab. (HINT: compounds containing the polyatomic ion carbonate and large metal cautions ARE NOT SOLUBLE AND WILL FORM A PRECIPITATE) *(4 points)*

   Reaction A_____
   _____
   Reaction B_____
   _____
   Reaction C_____
   _____
   Reaction D_____
   _____

3. Identify the types of reactions seen this lab. *(4 points)*

   Reaction A_____
   Reaction B_____
   Reaction C_____
   Reaction D_____

4. How do you know chemical changes occurred in each reaction? *(3 points)*
   _____
   _____

5. If you were to mass the reactants before the reaction and then mass the products after the reaction, what would you expect to find? Why? How is this related to balancing your equations? *(3 points)*

_____

_____

6. For each reaction type, explain using the 12 principles of green chemistry why you chose the reaction you did. *(8 points)*

	Part Chosen (1-8)	Criteria for choosing the procedure used	Green Chemistry Principle that guided your choice
Reaction A			
Reaction B			
Reaction C			
Reaction D			

Conclusion:
State whether or not the hypothesis was correct by referring to your observations from the lab. Be sure to match your theoretical products (what you predicted the products were) to the observed products (what was being formed) in each reaction. (5 points)

_____
_____
_____
_____

STAPLE THIS SHEET TO THE FRONT OF YOUR LAB AND HAND IT IN WITH YOUR LAB

KNOWLEDGE:	**Pre –Lab questions**	**/8**	**TOTAL:**	**/13**
	Materials	**/2**		
	Purpose	**/1**		
	Procedure	**/2**		
INQUIRY:	**Hypothesis**	**/2**	**TOTAL:**	**/30**
	Analysis questions	**/23**		
	Conclusion:	**/5**		
COMMUNICATION:	**Grammar, use of sci. vocab. and conventions.**	**/5**	**TOTAL:**	**/10**
	Observations	**/5**		

Student Pre-lab questions

For each of the following experimental procedures and observations;

a) Write a balanced chemical equation,
b) Name the product
c) Identify the type of reaction**.**

1. A student heats barium metal over a flame, and begins to react with oxygen gas in the air. A white, crystalline solid begins to form on the barium. (2 points)

2. A chemist mixes lead nitrate and potassium iodide solutions (both of which are clear liquids) and mixes them. A yellow precipitate appears in the final solution. The remaining liquid is an ionic solution. Solid precipitates containing lead are always yellow. (2 points)

3. Magnesium metal is placed it in sulfuric acid. The solution begins to bubble. The remaining liquid is an ionic solution. (2 points)

4. Water is placed in an electrolysis machine. The water placed in the machine begins to bubble. The gas is collected in two separate containers. It is noticed that when a glowing splint is placed in one gas, it flames up. In the other gas, a glowing splint produces a small popping sound signifying a tiny explosion. (2 points)

5. Analyze each reaction type in the procedure below. Of the two possible reactions, decide using the 12 principles of green chemistry which one of the two reactions your group will complete. Be prepared in the Discussion Section to explain why you chose these two reactions.

Properties of matter worksheet

Color:

Melting Point:

Boiling Point:

Solubility:

Electrical Conductivity:

Toxicity:

Impact on the environment:

……..

Properties of matter worksheet

Color:

Melting Point:

Boiling Point:

Solubility:

Electrical Conductivity:

Toxicity:

Impact on the environment:

ACTIVITY SERIES LAB (MICROSCALE)

Purpose: to study the chemical activity of common metals

Safety Precautions
Wear goggles at all times. Review precautions for handling acids. Since silver nitrate may stain skin and clothing, avoid contact.

Procedure (Part I):

1. Sketch out a matrix to record your data. Along the top of the matrix, you should list the six reagents you will be using in this part of the experiment: KNO_3, $Mg(NO_3)_2$, $Zn(NO_3)_2$, $CuSO_4$, $AgNO_3$, and distilled water. Along the side of the matrix, you should list the five metals used in this part of the experiment: copper, iron, magnesium, tin, and zinc.
2. Obtain a spot plate. In each of 6 wells put a small piece of copper. Add reagents according to the instructions below:

Well 1:	add 5 drops of KNO_3
Well 2:	add 5 drops of $Mg(NO_3)_2$
Well 3:	add 5 drops of $Zn(NO_3)_2$
Well 4:	add 5 drops of $CuSO_4$
Well 5:	add 5 drops of $AgNO_3$
Well 6:	add 5 drops of distilled water

3. Observe the wells for a few minutes. Record your observations.
4. Discard the contents of the wells <u>in the waste beaker</u>, making sure to return any unreacted metal.
5. Repeat step one using each of the following metals: magnesium, zinc, and tin.

Procedure (Part II):

1. Sketch out a matrix for this section of the experiment. There is one reagent (HCl) and five metals to be tested (copper, iron, magnesium, tin, and zinc).
2. To each of 5 wells, place one of the following metals: magnesium, zinc, tin, iron, and copper. To each of the wells, *carefully* add 5 drops of HCl. Observe the metals for several minutes and record your observations.
3. Discard the contents of the wells <u>in the waste beaker</u> and return any unreacted metal.

(over)

Questions (to be answered in complete sentences):

1. Which metal reacted with the most solutions?
2. Which metal reacted with the fewest solutions?
3. List the metals in order of their reactivity, starting with the most reactive metal. (The most reactive metal is the one that reacted with the most solutions, while the least reactive metal is the one that reacted with the fewest solutions). Such a ranking of elements is called an ***activity series.***
4. Based on your activity series, explain why the Statue of Liberty was made with copper instead of zinc.
5. Based on your activity series, which material might have been a better choice than copper for the Statue of Liberty? Why do you think it wasn't chosen?
6. Given your knowledge of relative chemical activity among these metals,
    a. Which metal is *most* likely to be found in an uncombined or "free" state in nature?
    b. Which metal would be *least* likely to be found uncombined with other elements?
7. Devise an experiment to investigate if Au is more reactive than Cu. Be specific. Use formulas in your answer.
8. Group 1A metals are more reactive than group 2A metals. Hence, you might suspect that elements become less reactive as atomic number increases in a period on the periodic table. Do the *transition metals* you studied in this experiment support this assumption? Be specific!

Lab Checklist:

- ❑ **Title Page (Title, Purpose, Name, Date, Period)**
- ❑ **Matrices for parts 1 and 2 (with detailed observations)**
- ❑ **Answers to questions <u>in complete sentences</u>**

Activity Series Lab – Observations for Part 1

	KNO_3	$Mg(NO_3)_2$	$Zn(NO_3)_2$	$CuSO_4$	$AgNO_3$	Distilled H_2O
Copper						
Iron						
Magnesium						
Tin						
Zinc						

Activity Series Lab – Observations for Part 2

	Reaction with HCl
Copper	
Iron	
Magnesium	
Tin	
Zinc	

Activity Series Chart

Metals

Non-Metals

Most Active

Name	Symbol	Name	Symbol
Lithium	Li	Fluorine	F
Potassium	K	Chlorine	Cl
Barium	Ba	Bromine	Br
Strontium	Sr	Iodine	I
Calcium	Ca		
Sodium	Na		
Magnesium	Mg		
Aluminum	Al		
Manganese	Mn		
Zinc	Zn		
Iron	Fe		
Cadmium	Cd		
Cobalt	Co		
Nickel	Ni		
Tin	Sn		
Lead	Pb		
Hydrogen	H		
Copper	Cu		
Silver	Ag		
Mercury	Hg		
Gold	Au		

Least Active

***

Elements **CANNOT** replace anything **ABOVE** them. The reaction **DOES NOT OCCUR** in this situation.

***

South Pasadena • Chemistry

Name_____
Period ___ Date ___/___/___

3 • What Happens When Chemicals Are Put Together?

PRECIPITATE LAB

1. For each compound, write the two ions that make it up. (Ex: ⑧ **K_2CrO_4** = K^+ CrO_4^{2-})
2. In each square, write the two new compounds that will form.
3. Cross out any compounds that you KNOW will be soluble in water and therefore will NOT be precipitates.

	⑧ **K_2CrO_4**	⑦ **$CuSO_4$**	⑥ **$Ba(OH)_2$**	⑤ **$Pb(NO_3)_2$**	④ **Na_2CO_3**	③ **KCl**	② **$(NH_4)_2SO_4$**
① **$AgNO_3$**	Ag_2CrO_4 ~~KNO_3~~						
② **$(NH_4)_2SO_4$**							
③ **KCl**							
④ **Na_2CO_3**							
⑤ **$Pb(NO_3)_2$**							
⑥ **$Ba(OH)_2$**							
⑦ **$CuSO_4$**							

Precipitate Lab

Procedure

1. Locate a Ziploc baggie containing a data table. Be certain that the baggie is clean and dry.
2. Place a drop of each of the two ionic compounds in each of the indicated boxes.

   Be careful not to touch the tip of the dropper bottle to the solutions. This will prevent contamination.
3. For each combination note whether there is a precipitate and its color (e.g. WHITE PPT).

   If no reaction occurs (the mixed drop remain clear) write N.R. for no reaction or draw a dash —.
4. If the precipitate is white, you may not be able to see it against the white background.

   Look at the drop sideways and/or slip a piece of colored paper below the droplets.

Data

	⑧ K_2CrO_4	⑦ $CuSO_4$	⑥ $Ba(OH)_2$	⑤ $Pb(NO_3)_2$	④ Na_2CO_3	③ KCl	② $(NH_4)_2SO_4$
① $AgNO_3$	red-brown PPT						
② $(NH_4)_2SO_4$							
③ KCl							
④ Na_2CO_3							
⑤ $Pb(NO_3)_2$							
⑥ $Ba(OH)_2$							
⑦ $CuSO_4$							

Clean Up

Rinse off the baggie with tap water, then with distilled water. (Tap water contains ions that may change the data.)
Dry the baggie with a paper towel.
If the baggie will not come clean, tell the instructor so he may replace the baggie.

Follow Up

1. Verify with the rest of the class which combinations result in precipitates and which do not. If you do not agree on the results, redo those questionable portions of the experiment. (Baggies that have not been cleaned properly or contaminated dropper bottles can affect the data.)
2. For the following combinations, write the molecular equation, the ionic equation, and cross out the spectator ions and write the net ionic equation.

1 & 8

- $2\ AgNO_{3\ (aq)} + K_2CrO_{4\ (aq)} \rightarrow Ag_2CrO_{4\ (s)} + 2\ KNO_{3\ (aq)}$
- $2\ Ag^+ + 2\ NO_3^- + 2\ K^+ + CrO_4^{2-} \rightarrow Ag_2CrO_{4\ (s)} + 2K^+ + 2\ NO_3^-$
- $2\ Ag^+ + CrO_4^{2-} \rightarrow Ag_2CrO_{4\ (s)}$

1 & 6

-
-
-

1 & 4

-
-
-

1 & 3

-
-
-

Table 4.1 Solubility Rules for Common Ionic Compounds

page 111

	Mainly water soluble
NO_3^-	all nitrates are soluble
CH_3COO^-	all acetates are soluble
Cl^-	all chlorides are soluble <u>except</u> AgCl, Hg_2Cl_2, and $PbCl_2$
Br^-	all bromides are soluble <u>except</u> AgBr, Hg_2Br_2,$PbBr_2$, and $HgBr_2$
I^-	all iodides are soluble <u>except</u> AgI, Hg_2I_2,PbI_2, and HgI_2
SO_4^{2-}	all sulfates are soluble <u>except</u> $CaSO_4$, $SrSO_4$,$BaSO_4$, $PbSO_4$, Hg_2SO_4,and Ag_2SO_4
	Mainly water insoluble
O^{2-}	all oxides are insoluble <u>except</u> those of the 1A elements
S^{2-}	all sulfides are insoluble <u>except</u> those of the 1A and 2A elements and $(NH_4)_2S$
CO_3^{2-}	all carbonates are insoluble <u>except</u> those of the 1A elements and $(NH_4)_2CO_3$
PO_4^{3-}	all phosphates are insoluble <u>except</u> those of the 1A elements and $(NH_4)_3PO_4$
OH^-	all hydroxides are insoluble <u>except</u> those of the 1A elements, NH_4OH, $Ba(OH)_2$, $Sr(OH)_2$, and $Ca(OH)_2$

Name: _____

Period: _____

Measure for Measure

Student Directions:
Follow the activity step by step as outlined below. Answer the questions for each part before going on to the next part. When instructed to check something with the instructor do this before going on to the next part.

Part I: Bittium and Tinium

1. Using a digital balance count out the number of bittium "atoms" which have a mass of 1 gram.

	number of "atoms"	mass
bittium		1 gram

2. Count out the same number of tinium "atoms" and determine their mass.

	number of "atoms"	mass
tinium		

a) Question: Explain why equal numbers of bittium and tinium "atoms" do not weigh the same amount.

_____

_____

b) Question: If you counted out enough bittium "atoms" to weigh one ton and then you counted out the same number of tinium "atoms", how much would the tinium "atoms" weigh?

_____

c) In general if you have the same number of bittium "atoms" and tinium "atoms" the tinium "atoms" will always be _____ times heavier.

This can be expressed as a ratio of their masses

$$\frac{\text{mass of tinium}}{\text{mass of bittium}} = \_\_\_\_\_$$

A. *Question: A school bus is 14 times heavier than a car. Express this as a mass ratio.*

$$\frac{\text{mass of school bus}}{\text{mass of car}} = \_\_\_\_\_$$

Explain what this means in words to your partner(s).

If the mass of one car is 1000 pounds, what is the mass of one bus? _____
(Does your answer make sense?)

B. *Question: A pen is 18 times heavier than a paper clip. What is the mass ratio?*

$$\frac{\text{mass of pens}}{\text{mass of paper clips}} = \_\_\_\_\_$$

If the mass of 50 pens is 1500 grams what is the mass of 50 paper clips? _____
(Does your answer make sense?)

3. a) Count out 40 bittium "atoms" and mass them.

	number of "atoms"	mass
bittium "atoms"	40	

b) Using the mass ratio determined before in part 2C, calculate what the mass of 40 tinium "atoms" should be. Show the calculation below in the space provided.

Hint: you can set up a ratio of tinium "atoms" to bittium "atoms" as shown below. Cross multiply and solve for X (the unknown which in this case is the tinium "atoms")

$$\frac{\text{tinium}}{\text{bittium}} = \frac{\text{tinium}}{\text{bittium}}$$

Calculated mass of tinium _____

c) After you have **calculated** this mass, actually count out 40 tinium "atoms" and see if they have the mass you predicted.

Actual measurement of tinium to "check" your calculation:

	number of "atoms"	mass
tinium	40	

Question: If you have the same number of tinium "atoms" as bittium "atoms", describe in words how you can always get the mass of one if you know the mass of the other. Discuss this with your partner(s).

_____

_____

_____

Part II: Bittium and Tinium

1. a) Obtain a film canister (IIA-tinium) with an unknown number of tinium "atoms". Mass the film canister full of tinium "atoms". Then mass an empty film canister. Determine the mass of the tinium "atoms" alone.

   Mass of film canister and tinium "atoms" _____

   Mass of film canister _____

   Mass of tinium "atoms" _____

b) If you had an equal number of bittium "atoms", how much would they weigh? Use the mass ratio between bittium and the tinium from 2C above to calculate the mass of an equal number of bittium "atoms". Show the calculation in the space provided below.

Calculated mass of bittium _____

c) Now, actually mass this many grams of bittium "atoms" and count them. _____

d) So, how many tinium "atoms" are in the film canister? _____ (check with instructor)
(Don't open the canister with the tinium!)

Part III: Goldium and Hexium

1. Mass a film canister (IIIA - hexium) containing an unknown number of hexium "atoms". Determine the mass of hexium.

   mass of film canister and hexium _____

   mass of film canister _____

   mass of hexium _____

   Mass a film canister (IIIA - goldium) which contains the same number of goldium "atoms". Determine the mass of goldium.

   mass of film canister and goldium _____

   mass of film canister _____

   mass of goldium _____

How much heavier is hexium than goldium? Can you express this as a mass ratio? (I bet you can!)

$$\frac{\text{mass of hexium}}{\text{mass of goldium}} = \_\_\_\_\_$$

Questions: Are the masses of hexium and goldium the same? _____

Are the number of "atoms" of hexium and goldium the same in each film canister? _____

Even though you have never seen what these *"atoms" look like or counted them for yourself, what do you know about hexium and goldium?* _____

_____

_____

3. Obtain a film canister (IIIB - hexium) containing an **unknown** number of hexium "atoms". Determine the mass of the hexium. (Don't open the canister!)

mass of film canister and hexium _____

mass of film canister _____

mass of hexium _____

If you had the **same number** of goldium "atoms", how much would they weigh? Use your mass ratio to calculate their mass in grams. Show the calculation in the space provided below.

Calculated mass of goldium _____

4. Now, actually mass this many grams of goldium "atoms" and count them.

	mass	number of "atoms"
goldium		

So, without opening the canister, how many hexium "atoms" were in IIIB-hexium? _____
(Check this answer with your instructor)

5. How many hexium "atoms" and goldium "atoms" were in the first two film canisters? THINK!! (Don't open the canisters!)

number of hexium "atoms" in canister IIIA - hexium _____

number of goldium "atoms" in canister IIIA - goldium _____

Part IV: Mysterium and Largium

1. Obtain a film canister (IVA - mysterium) containing an **unknown** number of mysterium "atoms". Determine the mass of the mysterium inside (without opening the canister).

Obtain a film canister (IVA - largium) containing the **same unknown** number of largium "atoms", and determine the mass of the largium (without opening the canister).

What is the mass ratio of mysterium to largium? _____

2. Obtain a film canister (IVB - largium) with a different unknown number of largium "atoms". Using the mass ratio of mysterium to largium, determine how much an equal number of mysterium "atoms" would weigh. _____

3. Get an open sample of loose mysterium "atoms". Do some measurements and calculations to figure out how many largium "atoms" are in IVB. _____ Check this answer with your instructor.

4. How many largium "atoms" were in the first film canister (IVA-largium)? _____

How many mysterium "atoms" were in the first film canister (IVA-mysterium)? _____

Moles of Chalk Lab

Weigh a piece of sidewalk chalk and then write your name on the sidewalk. Weigh the chalk again, and determine the number of moles of calcium carbonate that were used.

*Formula for calcium carbonate:*_____

*Weight of chalk before writing your name:*_____

*Weight of chalk after writing your name:*_____

*Grams of chalk required to write your name:*_____

Calculations (show all work):

I needed _____ *moles of calcium carbonate to write my name on the sidewalk.*

*Name:*_____

Name _____ Date _____

Penny Lab

The familiar Lincoln penny has been made of 95% copper from 1909-1982. However, the price of copper had risen so much that the U.S. treasury needed to change the composition of the penny. As of 1982, pennies have been made of mostly zinc, with a thin copper layer on the outside. In this lab, you will determine the percent zing and the percent copper in a penny. To do this, you will take advantage of the activity series of metals, and perform a single replacement reaction on the penny.

Procedure

Day 1

1. Obtain a clean, unscratched penny (post 1982, the newer the better). You and your lab partner will each need different years so you can tell them apart! Record the years in the data table.
2. Using a file, carefully scratch the edge of your penny in two places, until the zinc core is showing.
3. Weigh each penny. Record the mass.
4. Place each penny in a beaker with about 30 mL of HCl. Label the beaker with you names. Notice the bubbling around the filed area. Set beaker aside in designated area.

Day 2

5. Decant HCl and rinse penny with DI water. Carefully remove penny and dry with paper towel.
6. Once dry, weigh copper shell and record mass.

Calculations/Questions

1. How much zinc was lost in penny 1?

2. How much zinc was lost in penny 2?

3. What is the % Zn in penny 1?

4. What is the % Zn in penny 2?

5. What is the average % Zn?

	Penny 1	Penny 2
Name		
Penny Year		
Initial mass (g)		
Final mass (g)		

6. The actual % Zn in pennies is 97.6%. Calculate the percent error.

7. Write the complete, balanced equation for the chemical reaction that occurred.

Experiment 2: Determination of the Empirical Formula of Magnesium Oxide

GOAL AND OVERVIEW

The quantitative stoichiometric relationships governing mass and amount will be studied using the combustion reaction of magnesium metal. Magnesium is reacted with oxygen from the air in a crucible, and the masses before and after the oxidation are measured. The resulting masses are used to calculate the experimental empirical formula of magnesium oxide, which is then compared to the theoretical empirical formula. A crucible and Bunsen burner will be used to heat magnesium metal to burning.

Objectives of the data analysis:

- Determine the expected formula for the ionic oxide expected when Mg reacts with O_2
- Find the theoretical and actual yields of Mg_xO_y
- Evaluate results using stoichiometry and error analysis

SUGGESTED REVIEW AND EXTERNAL READING

- Data analysis introduction (pp. 13-19), reference section 3; textbook information on ionic compounds and empirical formulas

BACKGROUND

In 1778, Lavoisier concluded that combustion was a reaction of oxygen in the air with a sample of matter. He realized that as the substance burned gained mass, the same mass was lost from the surrounding air. A great deal of chemical knowledge has been amassed by using simple combustion experiments conducted with crucibles, burners, and balances. In this experiment, you are using this technique to experimentally determine the empirical formula of magnesium oxide.

This lab illustrates (i) the law of conservation of mass and (ii) the law of constant composition.

(i) The total mass of the products of a reaction must equal the total mass of the reactants
(ii) Any portion of a compound will have the same ratio of masses as the elements in the compound

Molecular composition can be expressed three ways:

(i) In terms of the number of each type of atom per molecule or per formula unit (the formula).
(ii) In terms of the mass of each element per mole of compound.
(iii) In terms of the mass of each element present to the total mass of the compound (mass percent).

The empirical formula of a compound gives the lowest whole-number ratio of the constituent atoms that is consistent with the mass ratios measured by experiment.

In this lab, magnesium metal (an element) is oxidized by oxygen gas to magnesium oxide (a compound). Magnesium reacts vigorously when heated in the presence of air. The Mg-O_2 reaction is energetic enough to allow some Mg to react with gaseous N_2. Although there is a higher percentage of N_2 gas in the atmosphere than O_2, O_2 is more reactive and the magnesium oxide forms in a greater amount than the nitride. The small amount of nitride that forms can be removed with the addition of water, which converts the nitride to magnesium hydroxide and ammonia gas. Heating the product again causes the loss of water and conversion of the hydroxide to the oxide.

The unbalanced equations are:

$$Mg_{(s)} + N_{2(g)} + O_{2(g)} \rightarrow MgO_{(s)} + Mg_3N_{2(s)} \quad (1)$$

$$MgO_{(s)} + Mg_3N_{2(s)} + H_2O_{(l)} \rightarrow MgO_{(s)} + Mg(OH)_{2(s)} + NH_{3(g)} \quad (2)$$

$$MgO_{(s)} + Mg(OH)_{2(s)} \rightarrow Mg_xO_{y(s)} + H_2O_{(g)} \quad (3)$$

Based on the masses of the solid reactant and product, the mass in grams and the amount in moles of Mg and O in the product can be determined:

mass of Mg + mass of O = mass of Mg_xO_y or mass of O = mass of Mg_xO_y – mass of Mg (4)

mol Mg = mass Mg/MM_{Mg} and mol O = mass O/MM_O (5a,b)

The empirical formula of magnesium oxide, Mg_xO_y, can be written based on the lowest whole-number ratio between the moles of Mg used and moles of O consumed.

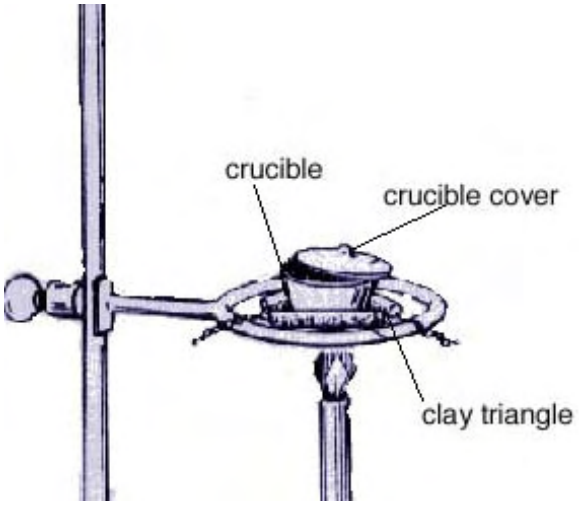
PRELAB HOMEWORK *(to be filled out in your bound lab notebook before you perform the experiment)*
Title and date
Define: (1) molecular formula; (2) empirical formula; (3) mass percent; (4) metal oxide;
Answer:

1. If the mass percent of each element in a compound is known, what computational steps are taken to determine the compound's empirical formula?
2. If the empirical formula of a compound is known, what additional information is required to determine the molecular formula of the compound?
3. Is the reaction of magnesium metal and oxygen gas an oxidation-reduction reaction? If so, what is the change in oxidation number of each type of atom?
4. What is the theoretical yield in grams of MgO if 2.54 g Mg metal reacts with excess O_2? What is the theoretical yield of Mg_3N_2 if the same amount of Mg reacts with excess N_2?

Procedure (Experimental plan)
Data tables

EXPERIMENTAL
Materials

- Safety goggles
- Magnesium ribbon, Mg
- Balance (to 0.0001g)
- Ring stand
- Bunsen burner
- Ring support/ clay triangle
- Crucible/ lid
- Tongs
- Clay tile



CAUTION:
Eye protection is essential.
Open flame will be present.
Do not breathe the fumes generated.
Once any burner is lit, assume ALL equipment is hot.
Do not touch the crucible, lid, triangle, ring, or stand during or after they have been heated.
Never place anything hot on a balance.

CRUCIBLE USE:

- Crucibles allow the heating of substances to high temperatures (like those encountered with burning metals) without risk of breakage.
- Do not touch the crucible with your hands (oils contaminate it and/or you could be severely burned).
- Do not place a hot crucible on a lab bench (the temperature difference may cause it to break).

Prior to starting:

- Practice using the tongs to pick up the lid from the crucible and the crucible from the clay triangle.
- Practice placing the lid partially over the crucible so that there is a gap of about 0.5 cm (the lid should rest on the crucible edge and two legs of the triangle).
- Practice placing the crucible with lid on the clay tile (when carrying the crucible, ***always*** hold it with tongs and support it with the tile).

EXPERIMENTAL PROCEDURE

Your TA will demonstrate. Ask questions as needed.

1) Fire the empty crucible and lid for about 3 minutes to remove water, oils, or other contaminants and to make sure there are no cracks. The bottom of the crucible should glow red-hot for about 20 seconds. Remove the flame and cool the crucible with lid.
2) Record the mass of crucible and lid once it has cooled. Do not handle it with your hands.
3) Obtain about 0.3 g (35 cm) magnesium ribbon (do not handle the ribbon with your hands). Fold the ribbon to fit into the bottom of the crucible.
4) Record the mass of the magnesium ribbon, lid and crucible.

5) Place the crucible securely on the clay triangle. Set the lid ***slightly*** off-center on the crucible to allow air to enter but to prevent the magnesium oxide from escaping.
6) Place the Bunsen burner under the crucible, light it, and brush the bottom of the crucible with the flame for about 1 minute; then, place the burner under the crucible and heat strongly.
7) Heat until all the magnesium turns into gray-white powder (probably around 10 minutes).
8) Stop heating and allow the crucible, lid and contents to cool.

9) Add about 1 ml (~10 drops) of distilled water directly to the solid powder. Carefully waft some of the gas that is generated toward your nose, but ***be very careful***. Record any odor.
10) Heat the crucible and contents, with the lid slightly ajar, ***gently*** for about 2 minutes and then strongly for about another 3 to 5 minutes.

11) Allow the crucible to cool and then record the mass of the crucible, lid and contents.
12) Follow instructions for oxide disposal given by your TA. Clean all equipment thoroughly.

DATA ANALYSIS

Report the following information and show sample calculations

1. mass of Mg metal used
2. theoretical yield of MgO from reaction: $Mg_{(s)} + \frac{1}{2}\ O_{2(g)} \rightarrow MgO_{(s)}$
3. mass of oxide product formed
4. mass of O incorporated (by difference; see eq. 4)
6. empirical formula of the oxide
7. percent by mass of Mg and O in the oxide
8. percent yield of $Mg + \frac{1}{2}\ O_2 \rightarrow MgO$ (actual yield/theoretical yield)×100%

REPORTING RESULTS – Complete your lab summary

If a report is required in place of a lab summary

Abstract

Results

Sample Calculations

- See list in data analysis section

Discussion/Conclusions

- How does your experimental empirical formula compare to the theoretical empirical formula – do they match?
- What are primary sources of error/deviation in the experiment?
- How would factors such as
  - (i) incomplete conversion of Mg_3N_2 to MgO or
  - (ii) residual $Mg(OH)_2$ in the product affect your results?
- Does this method appear to be a valid way to determine the formula of metal oxides?

Review Questions

REVIEW QUESTIONS

1. What evidence do you have that a chemical reaction took place?
2. If some of the magnesium oxide had escaped from the crucible as smoke during the reaction, would your mass percent calculation of magnesium be too high or too low? Explain.
3. If the surface of the Mg ribbon you used were covered with a thin oxide coating prior to the reaction, would your mass percent calculation of magnesium be too high or too low? Explain.
4. If you heated 0.3000 g Zn instead of Mg, what mass of oxide product would you expect to obtain?
5. Suggest a modification to the procedure that would be more likely to ensure that all the Mg would react completely with O_2.

STOICHIOMETRY

Stoichiometry concerns mass relations in chemical formulas and chemical reactions. In this experiment, you will investigate the stoichiometry of potassium chlorate. Potassium chlorate is a compound containing the elements potassium, chlorine, and oxygen. When heated strongly, the compound decomposes and all the oxygen can be driven off. A common way to determine if the gas is oxygen, is to collect the gas into a bottle. A glowing wood splint is then placed in the bottle. It will burst into flames in the presence of oxygen. The potassium and chlorine remain as the product potassium chloride, for which the formula is KCl. When potassium chlorate is heated, the mass loss is due to oxygen evolved from the potassium chlorate. The starting mass and mole ratios can be used to determine the theoretical yield. You can then determine the percent recovery. To speed the rate at which the decomposition occurs, manganese dioxide is used as a catalyst. The mass of the manganese dioxide remains constant. The reaction occurring is:

$$\text{Potassium chlorate} \xrightarrow[\text{heat}]{MnO_2} \text{Potassium chloride} + \text{oxygen gas}$$

PRE-LAB QUESTIONS

1. How do you determine the percent yield of a reaction?
2. If you react 2.500 g of potassium chlorate, how many grams of the potassium product will you get?
3. In a version of the lab you react 1.00 g of potassium chlorate and get 10.0 mL of oxygen. What is your percent yield?
4. In another version of the lab you react 5.00 g of potassium chlorate and get 2.00 grams of the potassium product. What is the percent yield?

PROCEDURE:

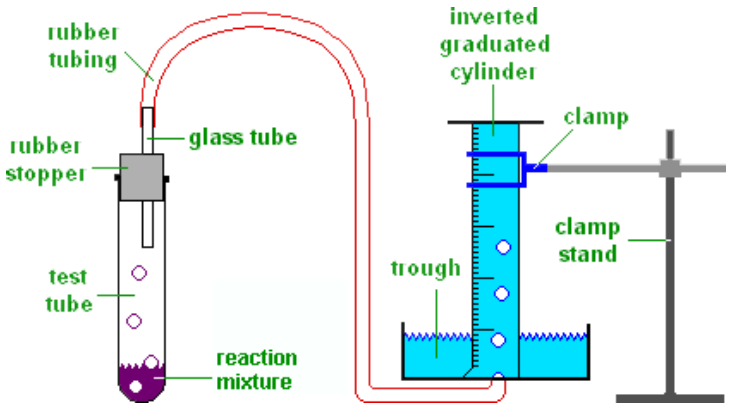
A. Determining Experimental Yield

1. To a clean, dry, small pyrex test tube add a pinch of manganese dioxide and weigh the tube plus manganese dioxide to the nearest 0.01 g on a laboratory balance.

2. Add about 0.75 g of dry potassium chlorate, and reweigh.

3. Tap gently to mix. Clamp the tube to a ring stand at an angle of about 45º.

4. Affix the gas collection apparatus as shown below.



5. Heat gently with a Bunsen burner flame.

6. When the solid melts, increase the heating and heat as strongly as possible for several minutes, until you can no longer see bubbles come up through the water. Allow the tube to cool and weigh.

7. Using stoichiometry calculate the theoretical moles and grams of KCl that should have been produced.

8. Calculate the percent yield of KCl.

9. Using stoichiometry, calculate the theoretical liters of oxygen that should have been produced.

10. Calculate the percent error for the oxygen gas collection.

B. Identification of product

1. After weighing the tube place a few mL of distilled water into the tube and dissolve some of the salt by stirring or capping and shaking the test tube.

2. Add a few drops of aqueous silver nitrate to the test tube. If you are told that the compound silver chloride is insoluble in water, does this test confirm the presence of chloride ion in the residue?

3. From the front of the room get a wood splint. Using your tongs hold one end of the wood splint in a Bunsen burner flame. When the wood splint starts to burn remove it from the flame. Carefully blow out any flame, but make sure that the wood is still glowing hot. Carefully remove the gas collection container from the water, being sure to keep the bottle upside down. Using your tongs place the glowing wood splint into the gas collection container. Record your observations.

NAME: _____
Lab Partner's Name: _____
DATE: _____

REPORT SHEET: STOICHIOMETRY

A. Determining Experimental Yield of potassium chlorate:

1. Mass of test tube and catalyst _____

2. Mass of test tube and catalyst and potassium chlorate _____

3. Mass of $KClO_3$ initial (line 2 less line 1) _____

4. Moles of $KClO_3$ Initial _____

5. Mass of test tube and catalyst and residue _____

6. Mass of oxygen lost (line 2 less line 5) _____

7. Mass of residue (line 5 less line 1) _____

8. Moles of KCl (residue) _____

9. Write and balance the equation for the reaction of $KClO_3$.

10. Convert Moles of $KClO_3$ (Line 4) to moles of KCl, Using the balanced equation. _____ mols

11. Theoretical yield of KCl. _____ g

12. Percent yield of KCl _____

$$\frac{\text{Experimental yield (line 7)}}{\text{Theoretical yield (line 11)}} \times 100 = \text{Percent Yield}$$

B. Identification of product

1. What anion in the residue can you test for?

2. What Reagent would you use to test for the above anion?

3. Write the balanced equation for the above reaction.

4. Record observations on what happened when you inserted a glowing splint into the gas collection tube. From these observations what was the gas collected in this reaction?

QUESTIONS: To accompany Stoichiometry of an Unknown Mixture

1. Balance the equation.

2. If you produce 3.63 moles of Na_2S, how many moles of H_2 were needed?

3. If 1.672 g of Na_2S is produced, how many grams of H_2 were needed?

Limiting Reactant Lab

In class, you've learned to compute how much of a chemical product you can make when you mix measured amounts of chemical reactants. In this lab, you will be actually using this information to predict how much product will be made; you will then calculate the percent yield gained from the amount that you actually recover.

In this lab, you'll be seeing the reaction of lead (II) nitrate with sodium bromide to form a lead (II) iodide precipitate and aqueous potassium nitrate.

Prelab:

1) If we mix 25 grams of sodium bromide with a large amount of potassium chloride, what will our theoretical yield of sodium chloride be?

2) If our actual yield from this reaction was 18 grams of sodium chloride, what would our percent yield for this reaction be?

3) Is the answer in question 2 reasonable? If so, explain why you think this was a reasonable answer. If not, explain what is wrong with it and discuss possible reasons you might get this answer in the laboratory.

4) What are some factors that might cause our percent yield to be greater than 100%? What are some factors that might cause it to be less than 100%? Make sure you discuss specific cases of how both might happen.

Materials:

Distilled water	100 mL beaker	250 mL beaker
2 weigh boats	Balance	100 mL graduated cylinder
Glass stirring rod	Sodium bromide	Lead (II) nitrate
Filter paper	Funnel	Forceps
Hot oven or incubator		

Procedure:

1) Pour 75 mL of distilled water into a 250 mL beaker using a graduated cylinder.

2) Using a weigh boat, weigh approximately 0.5 grams of sodium bromide on the balance. Make sure to record the exact amount of sodium bromide used, to the nearest 0.01 grams.

3) Add the potassium iodide into the distilled water. Stir until dissolved.

4) Using a clean weigh boat, weigh out approximately 1.0 grams of lead (II) nitrate. Make sure to record the exact amount of lead (II) nitrate used, to the nearest 0.01 grams.

5) Add the lead (II) nitrate to your beaker. Make sure to stir the mixture with a stirring rod to make sure that the reaction goes to completion. Let the reaction occur for approximately five minutes.

6) Obtain a piece of filter paper from your instructor and find its weight to the nearest 0.01 grams.

7) Fold the filter paper into quarters and place it in a funnel. Filter the solution through the filter paper, making sure you collect the filtrate in another beaker. This filtrate should be carefully labeled and set aside for later in the lab.

8) After you've finished filtering the solution, there may still be some chunks of lead (II) iodide stuck to the inside of the 250 mL beaker. Add a very small amount of distilled water to the beaker to rinse these into the funnel.

9) Once all of the lead (II) iodide has been collected in the filter paper, remove the filter paper from the funnel with forceps and press it between two dry paper towels. When most of the water has been removed, carefully place the filter paper in a hot oven. Let it sit in there for 15 minutes (or over night) to let the remaining water evaporate.

10) When the water has completely evaporated, find the mass of the filter paper and precipitate to the nearest 0.01 grams. Use this mass and the original mass of the filter paper to determine the weight of precipitate formed.

Analysis:

1) In this reaction, one of the reagents was a limiting reagent. Using your knowledge of chemical equations and limiting reagents, determine which was the limiting reagent for this experiment. Be sure to show your work!

2) When you determine the identity of the limiting reagent, determine the percent yield of lead (II) iodide formed.

3) Explain why your yield was not exactly 100%.

4) Devise improvements to your laboratory procedure/methods that would give you better results were you to do this experiment again.

Conclusion:

Comments:

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Ornament-Making

A Holiday Redox Activity

Introduction

Holiday fun—combine chemistry and art to design a holiday ornament!

Chemical Concepts

- Redox reactions
- Metal reactivity

Materials

Acidified copper(II) nitrate solution, $Cu(NO_3)_2$, 0.05 M, 25 mL
Galvanized (zinc-coated) iron, 2½″ × 2½″ piece
Hydrochloric acid solution, HCl, 6 M, 50 mL
Beaker, 150-mL
Beaker, 1-L
Cotton swab
Eraser
Graduated cylinder, 50-mL
Ornament hanger
Masking tape
Paper towels
Pencil
Scalpel
Tongs
Acrylic sealer (optional)

Safety Precautions

Hydrochloric acid solution is corrosive to skin and eyes and is moderately toxic by ingestion and inhalation. Acidified copper(II) nitrate solution is slightly toxic by ingestion and is a skin, eye, and mucous membrane irritant. The edges of the galvanized iron are sharp—be careful to avoid cuts and scratches. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Please review current Material Safety Data Sheet for additional safety, handling, and disposal information.

Preparation

1. Galvanized iron can be found at hardware stores—it is usually sold in large sheets. Cut into 2½″ × 2½″ squares using tin snips. Bore holes large enough to accept an ornament hanger into the pieces using either a drill or a drill press.
2. Prepare 250 mL of acidified copper(II) nitrate solution by dissolving 3 grams of copper(II) nitrate trihydrate, $Cu(NO_3)_2 \cdot 3H_2O$, in 250 mL of a 0.25 M hydrochloric acid solution.

Procedure

1. Completely cover both sides of a 2½″ × 2½″ piece of galvanized iron with masking tape. Mark where the hole is—this is the top of the ornament. Make sure that the edges of the galvanized iron are also covered.
2. Draw a simple design (such as your initials or name) on the masking tape with a pencil. Designs may be drawn on both sides of the piece of galvanized iron if desired. The design that is drawn will become the copper-colored part of the ornament. See Figure 1.
3. Use a scalpel to cut along the pencil marks. Remove the masking tape *inside* the drawing only so that the design is uncovered. See Figure 1.

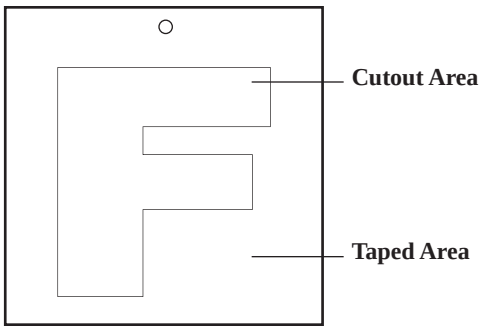

Figure 1.

4. Pour about 50 mL of 6 M hydrochloric acid solution into a 1-L beaker. Submerge the ornament in the hydrochloric acid solution using a pair of tongs. Lay the ornament flat on the bottom of the beaker so that it is completely submerged.
5. As soon as the rapid bubbling stops, remove the ornament from the hydrochloric acid solution using the tongs. Rinse the ornament with tap water and dry it with a paper towel.
6. Carefully clean the exposed area of the design by rubbing it with an eraser. Do not remove the masking tape.
7. Pour about 25 mL of copper(II) nitrate solution into a 150-mL beaker. Dip a cotton swab into the acidified copper(II) nitrate solution and gently rub the solution over the exposed parts of the design.
8. Once the entire design area is coated with copper, rinse the ornament with tap water and dry it with a paper towel. Remove the masking tape from both sides of the piece of iron. Attach a hanger to the hole in the top of the ornament.
9. *(Optional)* Coat both sides of the ornament with acrylic sealer. Hang the ornament so that it can dry completely.

Disposal

Please consult your current *Flinn Scientific Catalog/Reference Manual* for general guidelines and specific procedures governing the disposal of laboratory waste. Neutralize and dispose of the hydrochloric acid solution according to Flinn Suggested Disposal Method #24b. Dispose of the acidified copper(II) nitrate solution down the drain according to Flinn Suggested Disposal Method #26b.

Tips

- Explain to students that the places where the tape is removed will be the portion of the ornament that becomes copper-colored. Make sure that they design their ornaments accordingly. Suggest that the students keep the designs relatively simple, as simple designs tend to turn out best. Different designs can be drawn on each side of the ornament if desired. Make sure students note where the hole is located so that they orient their designs correctly. If they happen to prepare the ornament upside down or sideways, a new hole can be drilled in the top.
- Acrylic sealer may be applied to the ornaments to slow the tarnishing of the copper. Apply the acrylic sealer evenly to both sides of the ornament. Hang the ornaments in the laboratory and allow the sealer to dry completely.

Discussion

Galvanized iron is iron coated with a layer of zinc. The zinc protects the iron from rusting. The zinc layer is removed by submerging the galvanized iron in hydrochloric acid. The reaction between zinc and hydrochloric acid generates lots of bubbles from the formation of hydrogen gas.

$$Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$$

Once the zinc layer is removed, the iron surface then reacts with copper(II) nitrate. As the reaction occurs, a thin layer of copper metal is plated onto the surface of the iron.

$$Fe(s) + Cu(NO_3)_2(aq) \longrightarrow Fe(NO_3)_2(aq) + Cu(s)$$

In the "design areas" where the zinc layer has been removed, the underlying iron reacts with copper(II) ions, depositing copper on the surface and changing the color of the design area to reddish-orange. In areas where the galvanized iron does not undergo reaction with HCl, the ornament retains its original silver color.

Connecting to the National Standards

This laboratory activity relates to the following National Science Education Standards (1996):

Unifying Concepts and Processes: Grades K–12
Constancy, change, and measurement
Content Standards: Grades 9–12
Content Standard B: Physical Science, structure and properties of matter, chemical reactions

Reference

Fun with Chemistry; Institute for Chemical Education, University of Wisconsin—Madison, 1994.

The *Ornament-Making* activity is available as a student laboratory kit from Flinn Scientific, Inc.

Catalog No.	Description
AP5606	Ornament-Making Kit

Consult your *Flinn Scientific Catalog/Reference Manual* for current prices.

